Chapter 9
The Rate of Chemical Reactions

1. Chemical **kinetics** is the study of reaction rates.
2. The rate of the reaction is a measure of how fast a reaction occurs. It can be calculated as either:
   - The quantity of products produced per unit time OR
   - the quantity of reactants consumed per unit time.

CQD3. How is the rate of reaction defined?

A. Decrease in concentration of product
B. Decrease in concentration of product per unit time
C. Increase in concentration of products per unit time
D. Decrease in concentration of reactants per unit time

3. Different methods can be used to study the rate of a given reaction depending on the reaction itself. For example:
   - the volume of gas produced with time
   - the change in mass with time

CQD4. Calcium carbonate, $CaCO_3$ and hydrochloric acid, HCl react as shown below:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

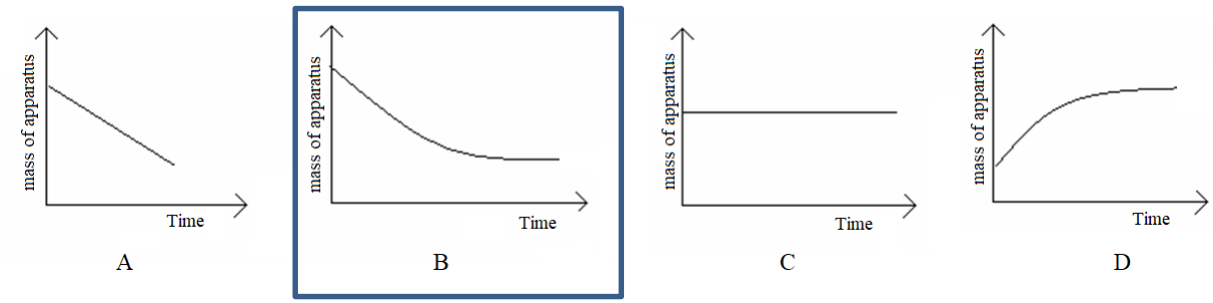
Choose the most suitable method to determine the rate of the reaction above?

A. Change in the temperature of the solution with time
B. Change in the volume of carbon dioxide gas with time
C. Change in the mass of water with time
D. Change in the concentration of hydrochloric acid with time

CQD5. Calcium carbonate, $CaCO_3$ and hydrochloric acid, HCl react as shown below:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

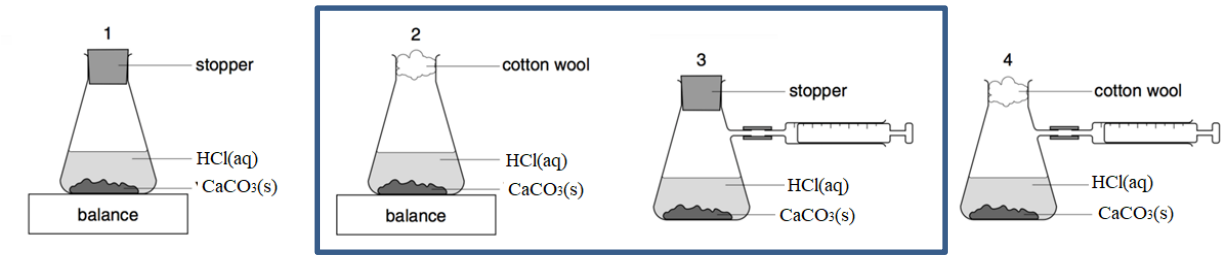
The mass of the apparatus used and its contents is plotted against time.
Choose the graph that shows what happens to the mass in time as calcium carbonate reacts with excess acid?



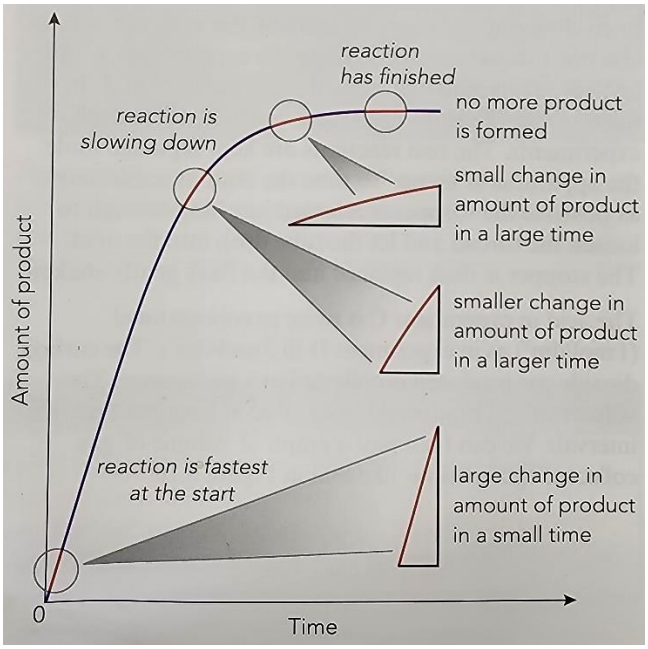
CQD6. Calcium carbonate, $CaCO_3$ and hydrochloric acid, HCl react as shown below:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Which diagram(s) would be suitable for studying the rate of this reaction?



4. Data is collected and plotted against time, for example the mass change versus time or the volume of gas produced versus time, …A typical plot is shown below.


reaction has finished
no more product is formed
reaction is slowing down
small change in amount of product in a large time
smaller change in amount of product in a larger time
reaction is fastest at the start
large change in amount of product in a small time
Amount of product
0
Time

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- The reaction rate is always the highest at the start, after that it starts decreasing with time. This happens because the reactants concentrations decrease as they get consumed.
- Eventually the amount of product no longer changes as the reaction stopped since at least one of reactants got consumed.
- The reaction rate at any instant can be calculated from the graph by drawing a tangent to the curve at the required time and finding its slope.

5. Collision Theory: For a reaction to occur

1) **particles of reactants must collide with one another**, as the number of particles present in a given volume increase, they will collide more often per unit time, that is the frequency of collisions will increases.
2) **they must collide with a minimum amount of energy called the activation energy**. Such collisions are called effective or successful collisions. As the kinetic energies f reactant particles increases, they will collide with higher energies. The number of collisions occurring with the activation energy or more per unit time will increase, hence the frequency of effective collisions will increase.
3) To increase the rate of a chemical reaction, it is required to increase the frequency of **effective** collisions, i.e increase the number of successful collisions per unit time.

6. Factors affecting rate of chemical reaction:

- nature of reactants (certain chemicals are more reactive than others by nature, the more reactive they are the faster they react)
- concentration of reactants or pressure if the reactant is a gaseous (as concentration of reactants increases, reaction rate increases)
- surface area of contact between reactants (or particle size in case of solids) in heterogenous reactions (reactions in which reactants exist in more than one state). As surface area of contact between reactants increases, reaction rate increases.
- temperature (as temperature increases, reaction rate increases)
- catalyst (addition of a catalyst increases reaction rate)

CQD1. Which of the following affects the rate of a chemical reaction?

a. concentration of reactants
b. pressure of gaseous reactants
c. surface of contact in case of heterogeneous systems
d. temperature
e. nature of substances reacting
f. catalysts
g. All of the above

CQD14. [G] What are the factors that affect the rate of a chemical reaction and how do they affect it?

- **Increasing the concentrations of the reactants increases the rate of the reaction.**
- **Increasing the temperature increases the rate of the reaction.**
- **Increasing the surface area of contact between reactants increases the rate of the reaction.**
- **Addition of a catalyst to a reaction increases the rate of the reaction.**

7. **Effect of changing concentration on rate of reaction:** Increasing the concentration of a reactant increases the number of particles in a given volume thus the reacting particles will collide more frequently so the number of collisions will increase per unit time, thus rate of reaction increases.

CQD18. [G] What is the effect of increasing the concentration of a reactant on the rate of a reaction? Justify your answer based on the collision theory.
Increasing the concentration of a reactant increases the rate of a reaction because the particles of reactants will collide more often which leads to more collisions.

CQD22. Consider two gases **A** and **B** in a container at room temperature. What effect will the following changes have on the rate of the reaction between these gases?

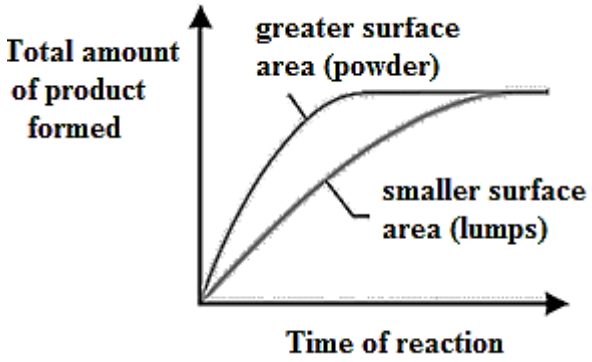
a. The pressure is doubled.
b. The number of molecules of gas *A* is doubled.
c. The temperature is decreased at constant volume.

a. To double the pressure means to either add more reactants or decrease the volume. Either act leads to an increase in the concentration. This means the frequency of collisions increases so the chances of successful collisions may increase leading to an increase in the rate.

b. Doubling the number of molecules of A, increases the frequency of collisions so the chances of successful collisions may increase leading to an increase in the rate.

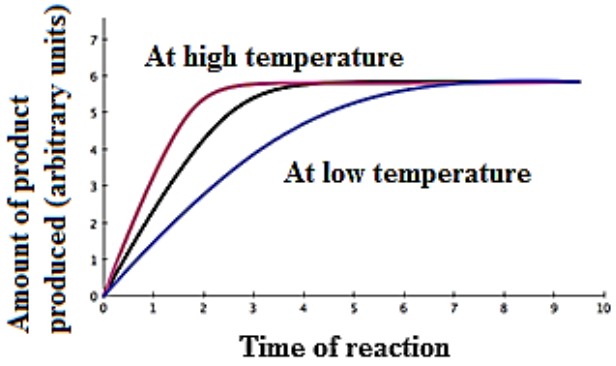
c. A decrease in the temperature decreases the average kinetic energies of the molecules which decreases their speeds. As a result, the frequency of effective collisions decreases which decreases the rate.

8. **Effect of changing pressure on rate of reaction:** if one or more of the reactants are gaseous, an increase in pressure can be achieved by either adding more of the gas in a constant volume or by decreasing the volume. Both of these events will lead to an increase in their concentration. Increasing the concentration, increases the number of particles in a given volume thus the reacting particles will collide more frequently so the number of collisions will increase per unit time, thus rate of reaction increases.
9. **Effect of changing surface area on rate of reaction:** if the surface area of contact between the reactants in a heterogenous system increases, then particles will collide more frequently and so the rate increases.



10. **Effect of changing temperature on rate of reaction**: increasing the temperature will increase the average kinetic energy of the reactant particles. As a result,
    - The frequency of collisions increases AND
    - the number of collisions occurring with the activation energy or more per unit time increases, in other words, the frequency of *effective* collisions increases

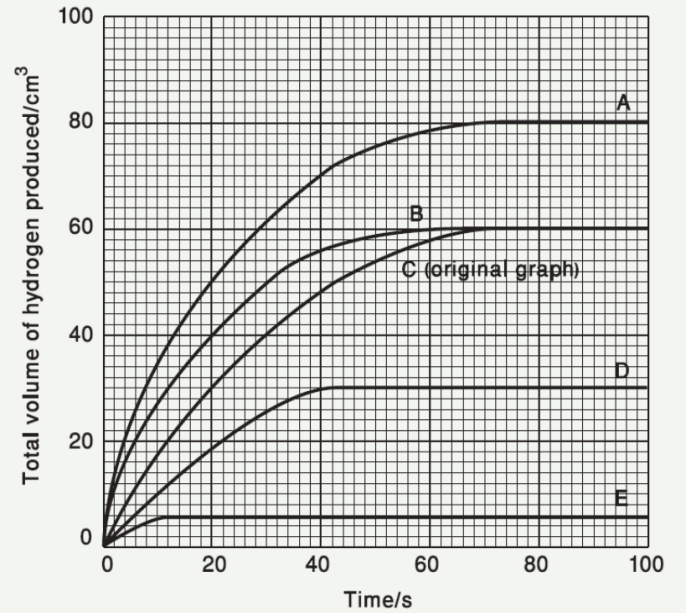
    Therefore, the reaction rate will increase.



CQD16. [G] What is the effect of increasing temperature on the rate of a reaction? Justify your answer based on the collision theory.

Increasing the temperature increases the rate of a reaction because the average kinetic energies of the reactants particles increase making them collide more often and more of the collisions occur with activation energy or more. In other words, the frequency of effective collisions increases.

CQD2. When magnesium ribbon reacted with excess dilute hydrochloric acid at room temperature, the total volume of hydrogen produced was noted every 20 seconds. The results obtained are given in the graph labeled C drawn below.



a) Which of the graphs A, B, D or E would have been obtained if:

i. A leakage of gas occurred during the experiment **E**

If there is a leakage then H_2 gas would escape, the collected volume will be much less.

ii. Only half the magnesium ribbon had been used as compared with the original amount used in C? **D**

Starting with half the mass of Mg results in the formation of half the volume of H_2, as production of hydrogen gas occurs in a 1:1 ratio with consumption of Mg

iii. Slightly warmer excess acid had been used? **B**

Slightly warming the mixture will increase the energy of the particles and their speeds. As a result the frequency of effective collisions will increase. This will increase the rate (decrease the time), shown by a steeper rise. However, since the same amount of Mg was used, then the same volume of H_2 will be produced at the end of the reaction..

iv. A greater mass of powdered magnesium had been used? **A**

A greater mass of Mg results in the formation of a greater volume of H_2. Powdering the Mg, increases the surface area exposed, so more collisions occur which leads to more effective collisions which leads to a faster rate.

b) What was the final volume obtained in experiment C? **60 cm^3**

SUMMARY

Change	Effect on rate of reaction	Explanation
Increasing the concentration	Increases	More particles per unit volume, frequency of collisions increases.
Increasing the temperature	Increases	More collisions occur with the activation energy or more, frequency of effective collisions increases.
Increasing the surface area	Increases	As surface area increases, more reactant particles are in contact frequency of collisions increases.

EX. 1 Which mixture will produce the slowest reaction to take place between solid barium carbonate chips and dilute hydrochloric acid?

a) low temperature ,large chips, and concentrated acid
b) Low temperatures, large chips and dilute acid
c) High temperature, small chips and dilute acid
d) high temperature , concentrated acid, and small chips
e) High temperatures, large chips, concentrated acid

EX. 2 Iron reacts with hydrochloric acid to form iron (II) chloride and hydrogen gas. 2g of iron is added to 100cm^3 of dilute hydrochloric acid. Which conditions produce the fastest reaction rate?

a) iron powder at 50ºC
b) iron pieces at 50 ºC
c) iron pieces at 25 ºC
d) iron powder at 25 ºC
e) iron powder at 35 ºC

EX. 3 For reactions to occur, the particles of the reactants must collide. Which of the following decreases the frequency of collisions between reactants particles?

a) Breaking the solid into smaller pieces in a reaction involving solid and liquid
b) Raising the temperature of the reaction mixture
c) Decreasing the concentration of a reactant
d) Increasing the pressure of the gases for reactions involving a gas
e) Decreasing the temperature of the reaction mixture

11. Health and safety issues: Increasing the surface area of solid (by reduction of particle size) may cause explosion in some cases. For eg.:

- In flour mills, the air can fill with fine flour dust particles which have a very large surface area relative to their volume. A spark will ignite the flour particles explosively.
- In coal mines where the air is filled with very fine coal dust and same can occur.

CQD13.

a. Explain why there is danger of explosion where a large amount of dry, powdered, combustible material is produced.

b. Explain (at the molecular level) why an increase in concentration of a reactant may cause an increase in rate of reaction.

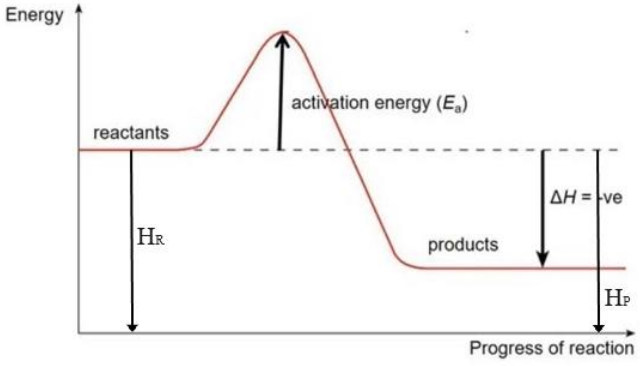
a. The large surface area provided by powdering, greatly increases the rate of the reaction. So, if sparked by any means, the reaction will be so fast and exothermic that an explosion results.
b. As the number of molecules present in a constant volume increases, it increases the possibility of collisions. When the frequency of collisions increases, the chances of successful collisions may increase. This leads to an increase in the reaction rate.

12. **Activation energy:** is the minimum amount of kinetic energy that must be overcome for chemical reaction to occur or it is the minimum KE that the reactants should collide with to react.

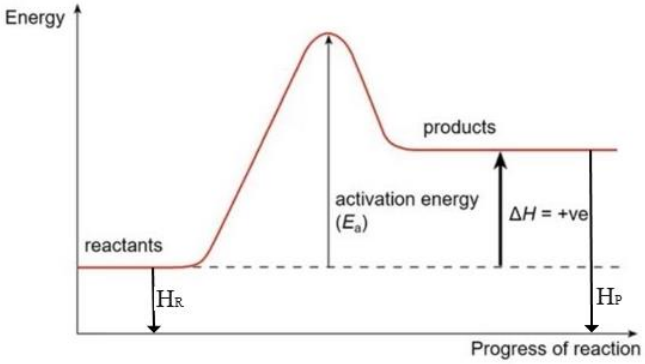
CQD21. Which of the following is the activation energy?

a) The maximum energy that the particles need to produce effective collision
b) The amount of energy used by the particles during a collision
c) **The minimum amount of energy required by reactants to undergo a reaction**
d) It is the minimum amount of energy which the reactant molecules must possess for the effective collision in forming the products.

13. Generally, reactions with low activation energy tend to be fast
14. Generally, reactions with high activation energy tend to be slow.
15. Reaction pathway diagrams labelled.

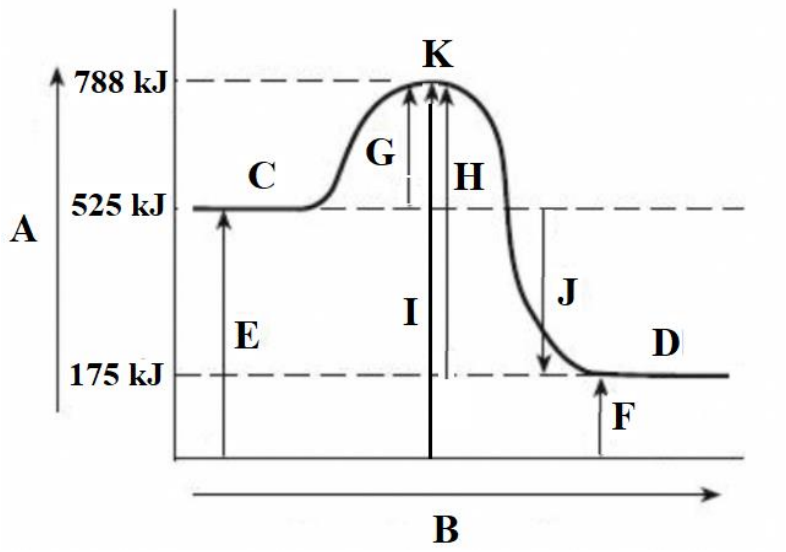


Exothermic Reaction



Endothermic Reaction

CQD19. Study the following potential energy diagram and answer the questions below.



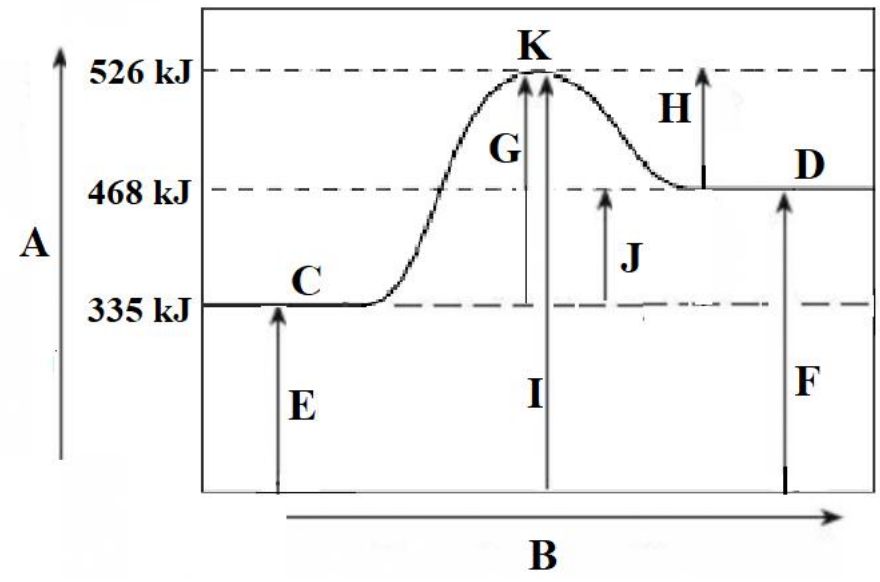
a. What do the labels A to K represent?
b. Calculate the values of E to J.
c. Is this an exothermic or endothermic reaction

a-b

- **A: Potential Energy**
- **B: Reaction coordinates**
- **C: Reactants**
- **D: Products**
- **E: $H_{reactants}$, Heat content of reactants (+525 kJ)**
- **F: $H_{products}$, Heat content of products (+175 kJ)**
- **G: E_{af}, Activation energy of the forward reaction (788 – 525 = +263 kJ)**
- **H: E_{ab}, Activation energy of the backward reaction (788 – 175 = +613 kJ)**
- **I: Energy of the activated complex (788 kJ)**
- **J: ΔH, Enthalpy change ($H_{products} – H_{reactants}$ = 175 - 525 = - 350 kJ)**
- K: activated complex

c. The overall reaction is **exothermic**; because **heat content of reactants is greater than heat content of products** and ΔH **< 0 (negative value)**

CQD20. Study the following potential energy diagram and answer the questions below.



a. What do the labels A to K represent?
b. Calculate the values of E to J.
c. Is this an exothermic or endothermic reaction

a-b

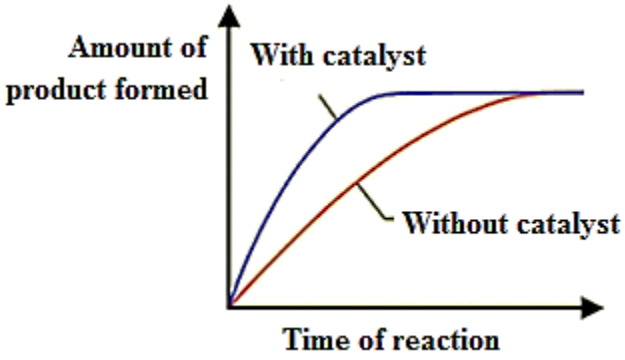
- **A: Potential Energy**
- **B: Reaction coordinates**
- **C: Reactants**
- **D: Products**
- **E: $H_{reactants,}$ Heat content of reactants (+335 kJ)**
- **F: $H_{products,}$ Heat content of products (+468 kJ)**
- **G: E_{af}, Activation energy of the forward reaction (526 – 335 = +191 kJ)**
- **H: E_{ab}, Activation energy of the backward reaction (526 – 468 = +58 kJ)**
- **I: Energy of the activated complex (526 kJ)**
- **J: ΔH, Enthalpy change ($H_{products} - H_{reactants}$ = 468 - 335 = + 133 kJ)**
- **K: activated complex**

c. The overall reaction is <u>endothermic</u>; because <u>heat content of reactants is less than heat content of products</u> and ΔH <u>> 0 (positive value)</u>

16. Catalyst is a substance that increases the reaction rate without itself being used up. It speeds up both forward and reverse reactions. It does NOT affect the reaction yield.

17. A catalyst takes part in the reaction, but is regenerated, so it is not used up.

18. A catalyst acts by providing a new path of lower activation energy for reactants to follow, so the catalyzed and uncatalyzed pathways proceed simultaneously.



EX. 4 A catalyst speeds up the reaction by

a) reacting with the reactants to form new products that do not need so much energy to form.
b) replacing the mechanism of the uncatalyzed reaction with a new mechanism that has a lower activation energy.
c) providing an additional path of a lower activation path so the catalyzed and uncatalyzed reactions proceed simultaneously.
d) slowing down the reverse reaction so the net result is a faster forward reaction.
e) providing energy to the reaction so the net activation energy becomes lower.

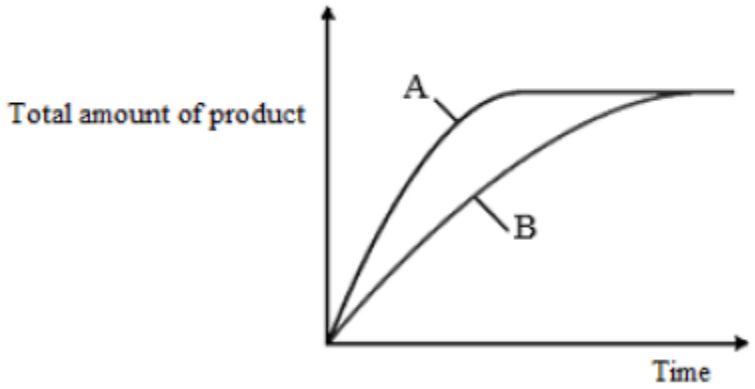
EX. 5 A catalyst speeds up

a) a reaction by lowering the activating energy of the forward reaction but increasing the activation energy for the reverse reaction.
b) both forward and reverse reactions by lowering ΔH.
c) the forward reaction without affecting the reverse reaction.
d) all reactions by lowering H reactant, H product and Δ H.
e) both forward and reverse reactions by lowering the activation energy of the forward and backward reactions.

19. Almost all reactions occurring in living cells require catalyst. These are proteins synthesized by cells called enzymes. Enzymes are biological catalysts, ie., catalysts made by living cells. General characteristics of enzymes include:

- Very efficient, only small amounts needed
- Very specific, each works only on particular reactions or on a specific type of molecules.
- They are sensitive to temperature; they can be destroyed completely (denatured) at temperatures above 45°C
- They are sensitive to acidity

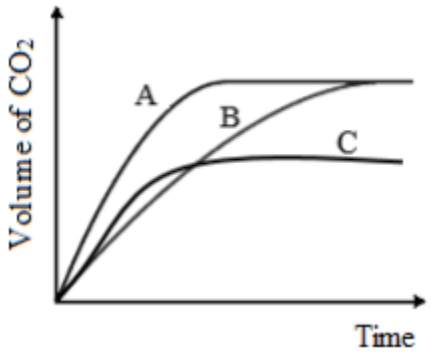
CQD17. The graph below shows the differences in the rate of a reaction at different conditions.



Indicate which of the following curves (A or B) is associated with:
a) a lower temperature. _____
b) a higher concentration. _____
c) a greater rate of a reaction. _____
d) a smaller surface area. _____
e) the addition of a catalyst. _____

a) B
b) A
c) A
d) B
e) A

CQD15. [G] In three different experiments, equal masses of calcium carbonate were reacted with the same volume of hydrochloric acid. In each experiment one factor affecting the reaction rate was changed. The volume of carbon dioxide released was measured every 5 seconds and the following graphs show the results.

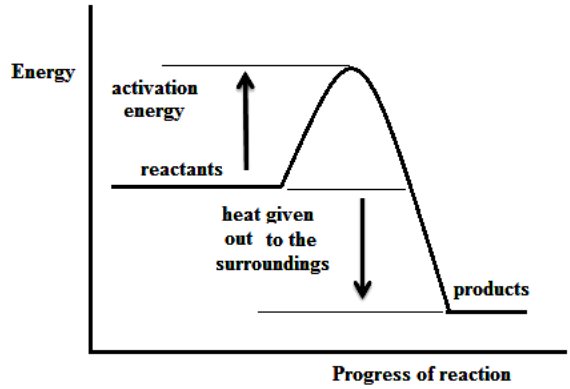
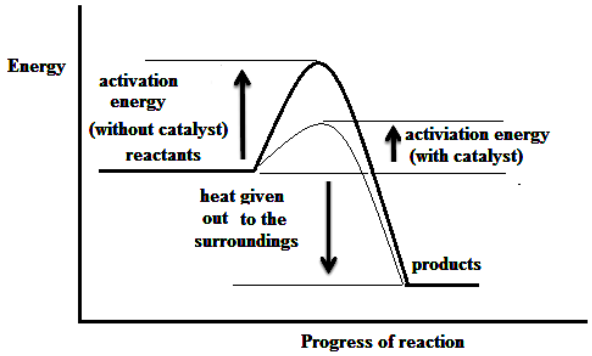
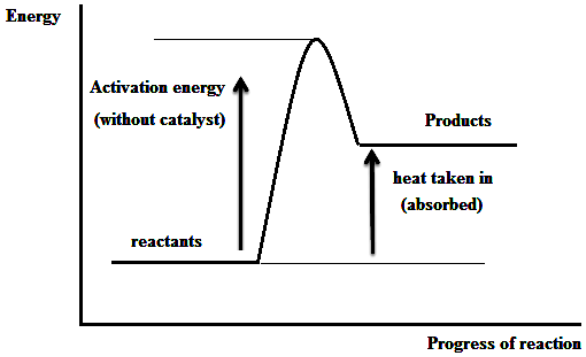
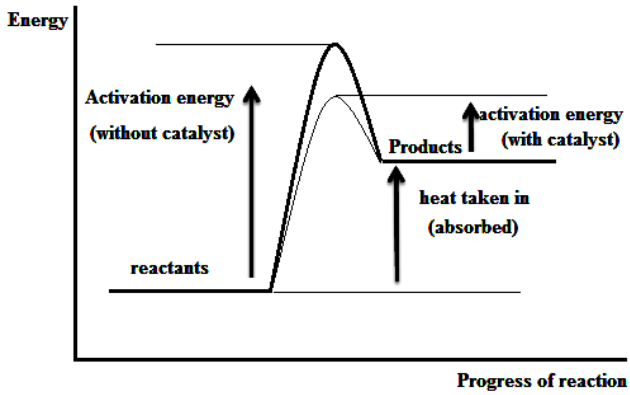


Which of the above curves can be associated with

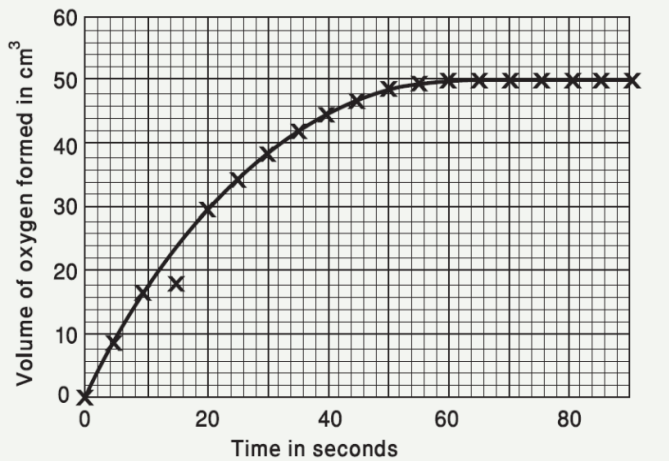
a. a lower temperature? _____
b. a lower concentration of hydrochloric acid? _____
c. a smaller surface area of calcium carbonate? _____
d. the addition of a catalyst? _____

a. B
b. C
c. B
d. A

20. Effect of catalysts on reaction pathway diagrams.

Exothermic Reaction	Endothermic Reaction
• **Without a catalyst**  • **Using a catalyst**  	• **Without a catalyst**  • **Using a catalyst** 

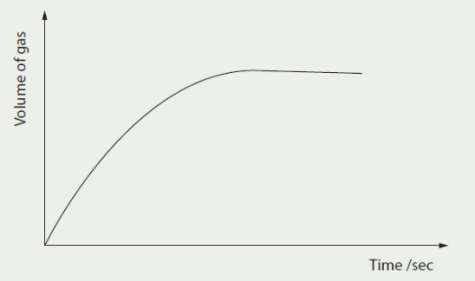
CQD7. The graph below shows the volume of oxygen formed during the decomposition of 50 cm^3 of hydrogen peroxide solution using 1 g of the catalyst.



a. Give one possible reason why the point (X) at 15 seconds is not on the graph line.
b. What volume of oxygen is collected after 20 seconds?
c. How long does it take to collect 40 cm^3 of oxygen?
d. When the reaction stopped, 1 g of the catalyst was still left. What caused the reaction to stop?

a. Experimental error such as wrong reading of the volume.
b. The volume of oxygen collected after 20 seconds is 30 cm^3.
c. It takes 33 seconds to collect 40 cm^3 of oxygen.
d. The reaction stopped because the reactants were totally consumed, all the hydrogen peroxide decomposed completely.

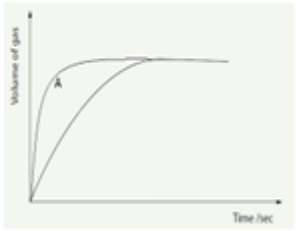
CQD8. A student added 10 g of powdered $CaCO_3$ to 150 cm^3 of 2 M acetic acid solution (an excess) at 60°C. CO_2 gas was produced and collected in a syringe. The volume of the gas was measured at regular time intervals. The results obtained are shown below.



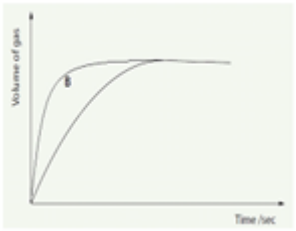
Sketch on the same set of axes the shape of the curve you expect to get if the following changes were made. Assume that one change was made at a time.

a. 300 cm^3 of 4 M ethanoic acid was used. Label your graph A.
b. Sulfuric acid, of the same volume and concentration was used. Label your graph B.
c. The acid was cooled before the reaction. Label your graph C.
d. 10 g of $CaCO_3$ in the form of small pieces were used. Label your graph D.
e. A catalyst was used. Label your graph E.
f. 12 g powdered $CaCO_3$ was used (the acid is still in excess). Label your graph F.

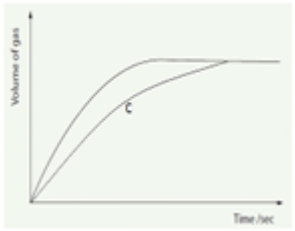
a. A larger concentration of acid increases the rate of the reaction. The graph will have a steeper start. The starting amount of $CaCO_3$ did not change so the graph will have the same ending as the original.



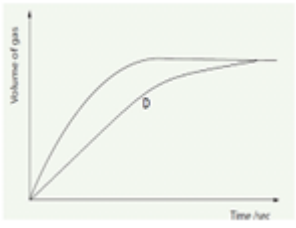
b. A stronger acid is used which increases the rate of the reaction. The graph will have a steeper start. The starting amount of $CaCO_3$ did not change so the graph will have the same ending as the original.



c. Cooling the acid decreases the rate. The particles speed and energy decrease with a drop in temperature. This decreases the number of effective collisions which decreases the rate. The graph will have a less steep start. The starting amount of $CaCO_3$ did not change so the graph will have the same ending as the original.



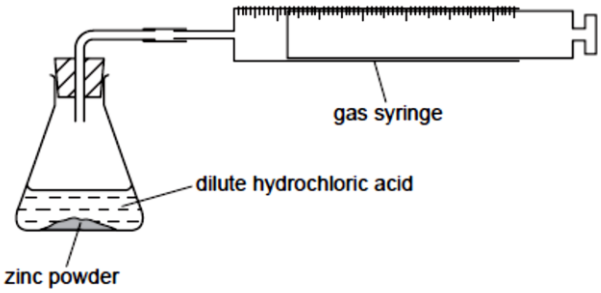
d. A decrease in the surface area decreases the possibility of collisions. As a result the rate will decrease. The graph will have a less steep start. The starting amount of $CaCO_3$ did not change so the graph will have the same ending as the original.



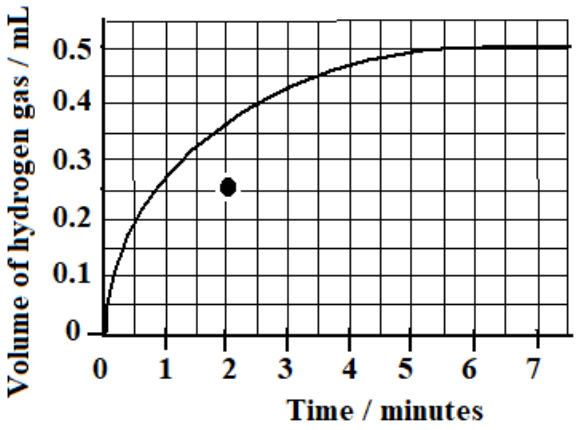
e. By definition a catalyst increases the rate of a reaction therefore graph will have a steeper start. The starting amount of $CaCO_3$ did not change so the graph will have the same ending as the original.

f. Increasing the starting quantity of solid will not affect the rate. However, the volume of gas produced will increase. The graph will have the same steepness but will end at a higher volume of gas.

CQD9. A student carried out the following reaction to study the speed of reaction between zinc and hydrochloric acid. Excess zinc powder was reacted with 50 cm^3 of dilute hydrochloric acid. The experiment was carried out at 25°C.



The volume of the gas was measured every minute for six minutes and the graph below was drawn.
Using the graph below answer the questions.



a. At what time does the result appear to be inaccurate? Explain
b. How does the speed of the reaction change over the six minutes? Explain
c. Sketch, on the grid, the graph you would expect if the experiment was repeated:
- at 50°C
- using excess lumps of zinc

d. When did the reaction finish? Why?
e. Explain, using the collision theory, why the reaction rate decreases.
f. What is the purpose of the gas syringe?

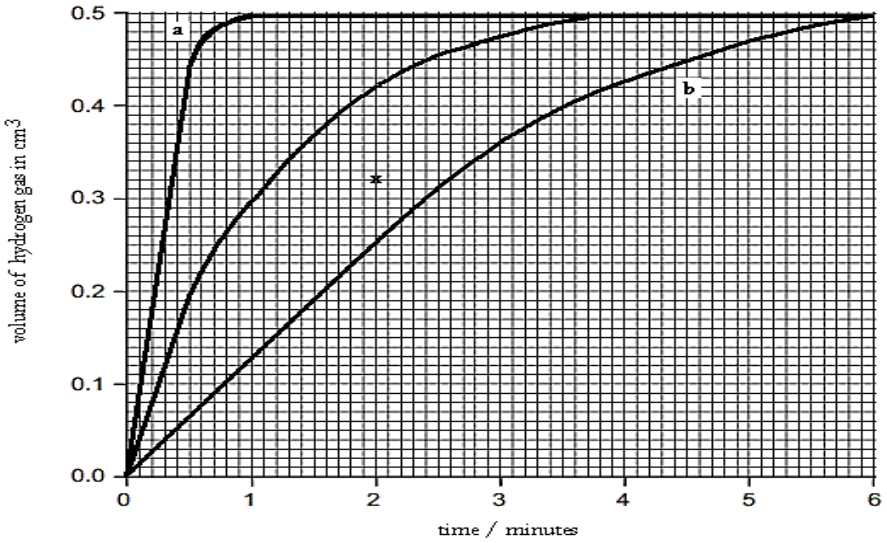
a. The result at 2 minutes. This result does not fit to the graph.

b. The speed decreases or slows down until it stops.

This occurs because the hydrochloric acid is being used up or becoming less concentrated.

(Never say the reactant is used up, you must specify which reactant)

c.



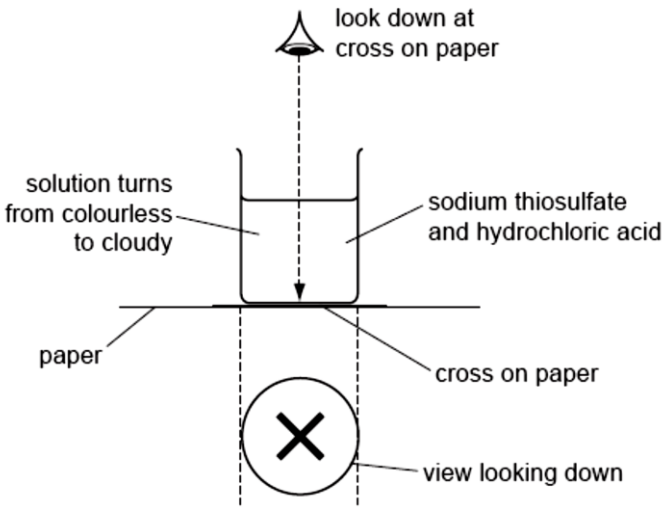
d. The reaction finished after 3.7 minutes. <u>All</u> the hydrochloric acid was used up.

e. As time passes, the initial concentration of hydrochloric acid decreases, number of particles decreases, thus the frequency of collisions decreases.

f. To measure the volume of the gas produced accurately.

CQD10. A student carried out the following experiment to investigate the speed of the reaction between sodium thiosulphate and dilute hydrochloric acid.

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + H_2O + SO_2$$



A 100 cm^3 beaker containing 50 cm^3 M sodium thiosulphate was placed on a black cross. 5.0 cm^3, 2.0 M hydrochloric acid were added and the clock was started. The experiment was carried out at 20°C.

Initially the cross was clearly visible. When the solution became cloudy and cross could no longer be seen, the clock was stopped and the time was recorded.

The experiment was repeated using 25 cm^3 0.2 M sodium thiosulfate and 25 cm^3 water. The results of this experiment and other experiments were recorded in the table below.

experiment	1	2	3	4
volume of thiosulfate / cm^3	50	40	25	10
volume of water / cm^3	0	10	25	40
volume of acid / cm^3	5	5	5	5
total volume / cm^3	55	55	55	55
time / s	48	60	96	.....

a. Why is it necessary to keep the total volume the same in all experiments?
b. If the experiment is performed at 42°C, explain using the collision theory what happens to the rate of the reaction.
c. Why does the cross on the paper disappear?
d. Experiment was repeated using 250 cm^3 beaker instead of the 100 cm^3 beaker. How would the results differ? Explain.

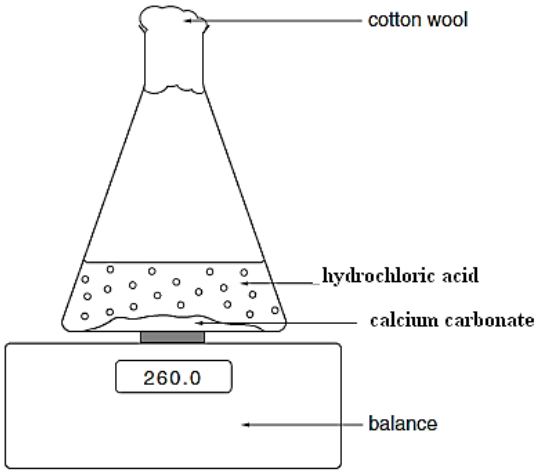
a. For fair comparison
The concentration of sodium thiosulfate is proportional to the volume of the solution.

b. The rate of reaction increases. As temperature increases, particles gain kinetic energy and move faster, thus the frequency of effective collisions increases.

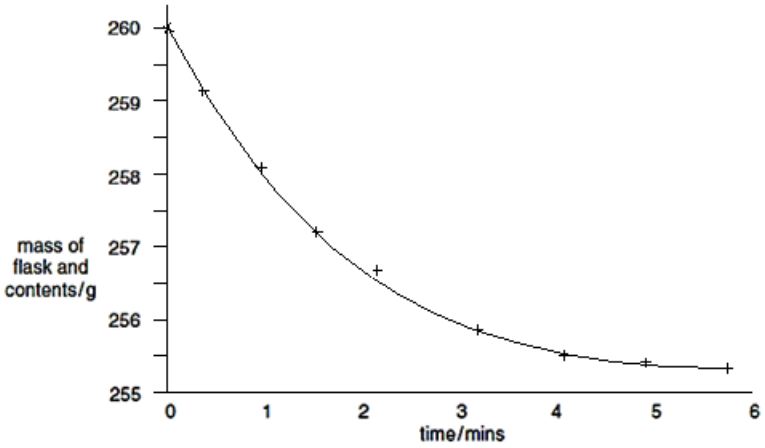
c. Due to the formation of sulfur precipitate.

d. The time taken for the cross to disappear would be longer. The bottom side of a larger beaker has a larger surface, so the precipitate formed spreads out. As a result, longer time is needed for enough of the precipitate to sediment and cover the cross.

CQD11. Dilute hydrochloric acid was added to excess calcium carbonate at 25°C and a gas was produced.

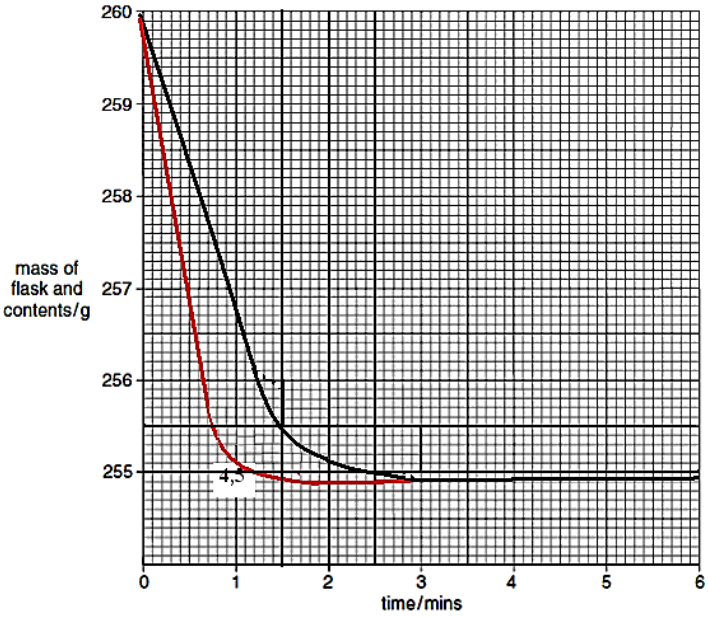


The flask was placed on the balance and the mass of the flask and its contents were recorded every minute. The following graph was plotted.

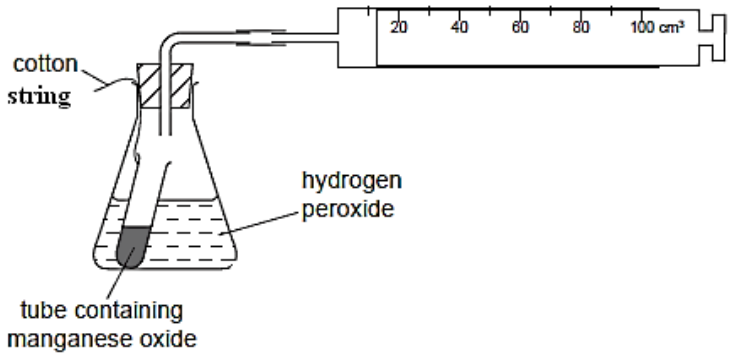


a. Name the gas produced**.**
b. What is the role of cotton wool.
c. Why does the mass of the beaker decrease?
d. On the same grid, sketch the graph if excess calcium carbonate powder was used at 50°C.
e. On the same grid, sketch the graph is concentrated hydrochloric acid was used instead.

a. Carbon dioxide
b. It allows the gas to escape while preventing the splashing of the acid.
c. Due to the release of a gas.
d. Refer to the graph 4
e. Refer to the graph 5



CQD12. The following apparatus is used to prepare oxygen. A tube of 1.0 g of powdered manganese oxide was added to the hydrogen peroxide solution.



a. Why was the tube of manganese oxide suspended in the flask?
b. What is the role of manganese oxide? Explain.
c. If 1.0 g lumps of manganese oxide were used, what is the effect on the result? Explain.
d. When the reaction stopped, 1.0 g of catalyst was still left. Why did the reaction stop?
e. How does the reaction start?

a. To prevent mixing of hydrogen peroxide and manganese oxide
or
So hydrogen peroxide and manganese oxide don't react.

b. Role: It acts as a catalyst, it speeds up the rate of the reaction.
Explanation: a catalyst speeds up the rate of reaction by offering an alternative pathway with lower activation energy for the reactant to follow.

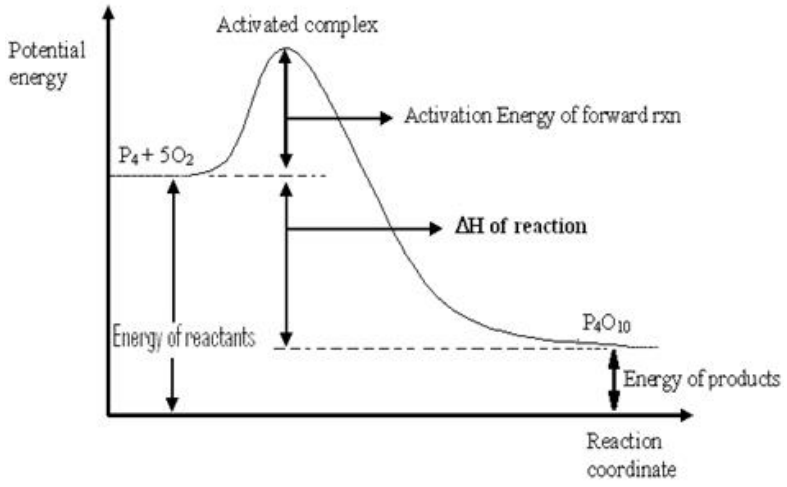
c. Effect: slower rate of reaction.
Explanation:1.0 g lumps have a smaller surface area than the powdered form, thus slower reaction rate.

d. All the hydrogen peroxide was completely decomposed.

e. By shaking the flask or letting go of the cotton string.

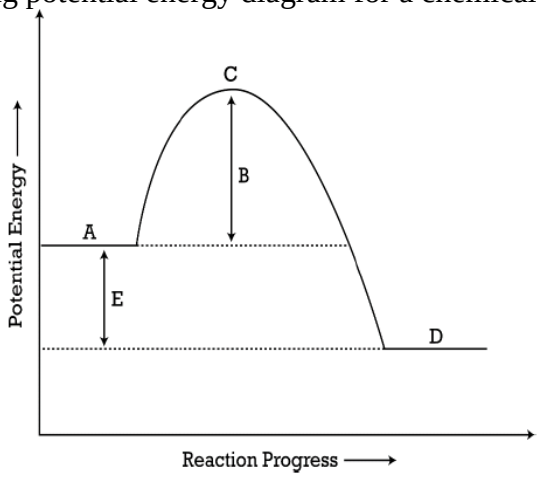
CQD23. Phosphorus, P_4, exposed to air burns spontaneously to give P_4O_{10}. The ΔH of this reaction is – 12321 kJ/mol P_4.
a. Draw an energy diagram for the net reaction, explaining the critical parts of the curve.
b. How much heat is produced when 12.4 g of phosphorus burn?

a.



b. 12321 kJ are released when 124 g of P_4 burn
? kJ are released when 12.4 g of P_4 burn $\Rightarrow$ 1232 kJ

CQD24. [G] Given the following potential energy diagram for a chemical reaction.



a. Identify the labels A, B, C, D, and E.
b. Give the effect of adding a suitable catalyst on each of the values B and E.

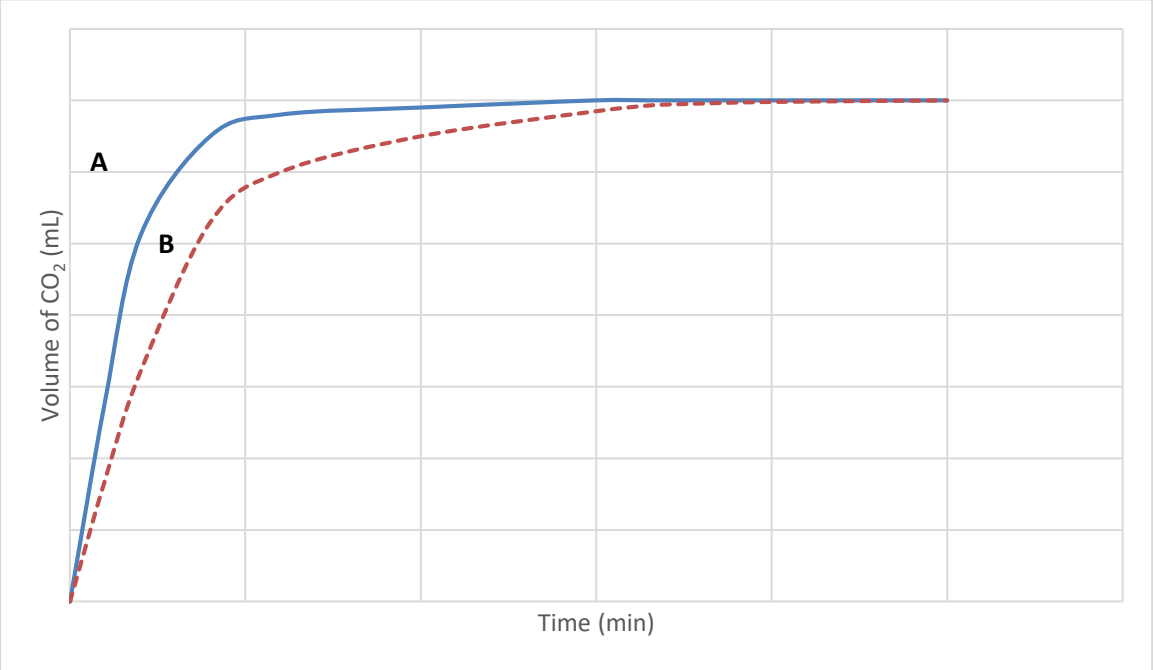
a. A= reactants or potential energy of reactants
B= activation energy of forward reaction
C= activated complex
D= products or potential energy of products
E= enthalpy change of the reaction

b. B decreases while E stays the same.

CQD27. [T] Calcium carbonate reacts with excess hydrochloric acid to produce water, carbon dioxide and a salt.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

The rate of the reaction is determined by measuring the volume of carbon dioxide gas released over time. The graph below shows the results of two experiments, **A** and **B**.



Which of the following best describes the change of condition between the two experiments?

A. The temperature is higher in **B** than it is in **A**.
B. Granulated calcium carbonate is used, in B, instead of powdered sample in A.
C. A catalyst is added in **B**.
D. The mass of calcium carbonate consumed in **A** is greater than that in **B**.
E. The mass of calcium carbonate consumed in **B** is greater than that in **A**.

CQD28. [T] The rate of the following reaction can be increased by increasing the concentration of the hydrochloric acid solution. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
Which of the following statements is/are true about the reacting particles?

1. The collision rate of particles stays the same.
2. The proportion of particles with sufficient energy to react decreases.
3. The collision rate of particles increases.
4. The proportion of particles with sufficient energy to react increases.

A. 1 only
B. 2 only
C. 3 only
D. 1 and 2 only
E. 1 and 4 only

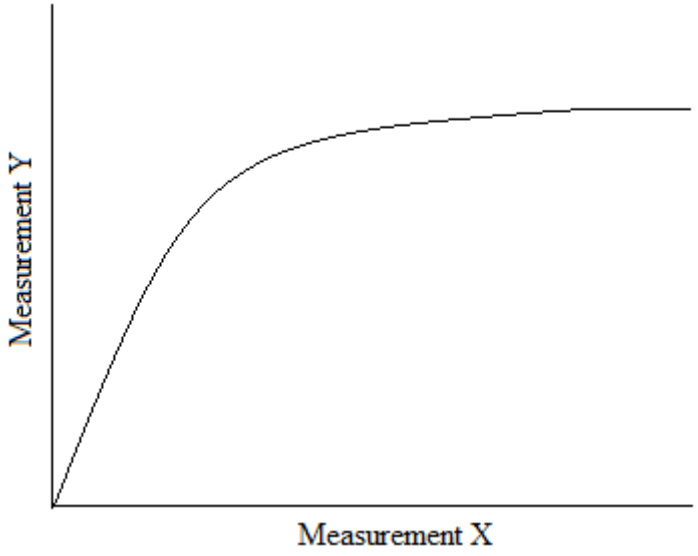
CQD29. [T] How does a decrease in temperature decrease the rate of a reaction according to the collision theory?

	collision rate of particles	proportion of particles with sufficient energy to react
A	decreases	stays the same
B	**decreases**	**decreases**
C	stays the same	decreases
D	stays the same	stays the same

CQD30. [T] A student investigates the rate of reaction between excess magnesium and dilute hydrochloric acid. The setup of his experiment includes a syringe.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

a. The student uses the collected data and draws the graph shown, in order to describe the rate of the reaction.



i. Identify measurements X and Y taken by the students and used for the graph.
ii. How does the rate of the reaction change as the reaction proceeds? Explain your answer.

b. Another student performs the same experiment at lower temperature.
i. What is the impact of this change on the rate of the reaction?
ii. Explain the change in rate in terms of collisions.

c. What is the effect of using smaller pieces of magnesium on the rate of the reaction? Explain your answer.

a. i. Measurement X is time, measurement Y is volume of gas released
ii. rate decreases / reaction gets slower because concentration of acid decreases leading to fewer collisions per unit time

b. i. The rate of the reaction decreases.
ii. Particles have less kinetic energy / particles move slower less the particles have sufficient energy to react / less probability of effective collisions

c. Higher rate due to larger surface area / more collisions between reactants

CQD31. [T]
A student investigated the effect of temperature change on the rate of reaction between dilute hydrochloric acid and sodium thiosulfate.

When these substances react they form yellow sulfur. The time taken for sulfur to be formed using same solutions but at different temperatures is recorded.

Six experiments were done.

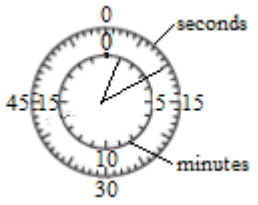
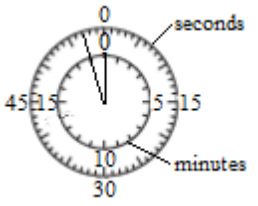
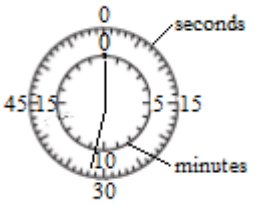
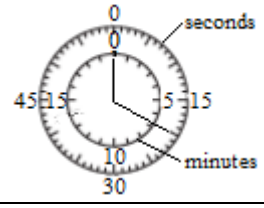
Experiment 1
- A paper with a marked cross was placed under a conical flask containing 10 cm^3 of sodium thiosulfate solution, so that the cross is visible through the solution from above.
- 10 cm^3 of dilute hydrochloric acid was poured from a measuring cylinder into the container of sodium thiosulfate solution and a chronometer was started at the same time.
- The temperature of the reaction mixture was measured and recorded.
- The chronometer was stopped when the cross mark cannot be seen anymore due to the formation of yellow sulfur.
- The final temperature of the solution was measured and recorded.
- The conical flask was emptied and rinsed with distilled water.

Experiment 2
- A paper with a marked cross was placed under a conical flask containing 10 cm^3 sodium thiosulfate solution, so that the cross is visible through the solution from above.
- 10 cm^3 of dilute hydrochloric acid was poured from a measuring cylinder into the container of sodium thiosulfate solution and a chronometer was started at the same time.
- The reaction mixture was heated to 30 °C.
- The chronometer was stopped when the cross mark cannot be seen anymore due to the formation of yellow sulfur.
- The final temperature of the solution was measured and recorded.
- The conical flask was emptied and rinsed with distilled water.

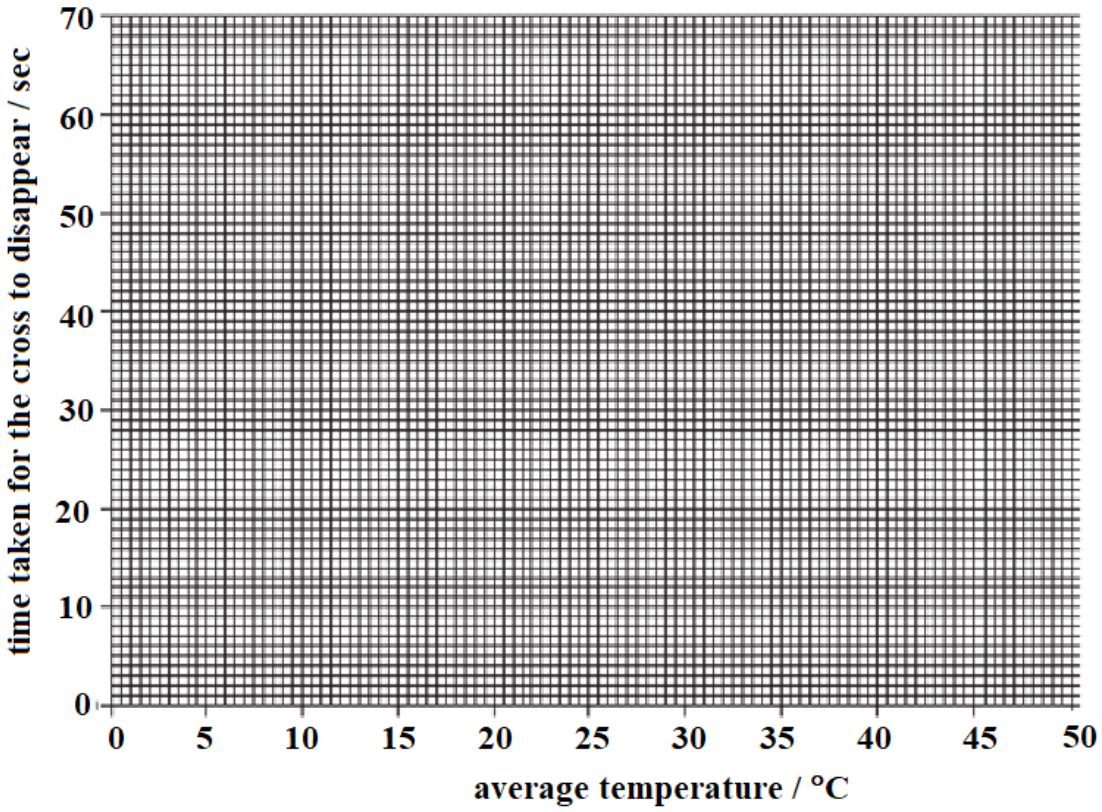
Experiment 3
The student repeated the procedure followed in experiment 2 but the reaction mixture was heated to 45°C.

Experiment 4
The student repeated the procedure followed in experiment 2 but the reaction mixture was heated to 50°C.

a. Complete the data table below by calculating and recording the average temperature and recording the time taken for the cross on the white paper to become invisible.

experiment number	initial temperature / °C	final temperature / °C	average temperature / °C	chronometer diagram	time taken for the cross on the white paper to become invisible / sec
1	17	21			
2	30	28			
3	45	43			
4	50	48			

b. Draw a smooth line graph on the grid below plotting results for experiments 1-4.



c. Use your graph to deduce the average temperature needed for the cross mark not to be seen after 60 sec. Show on the grid how you found your answer.

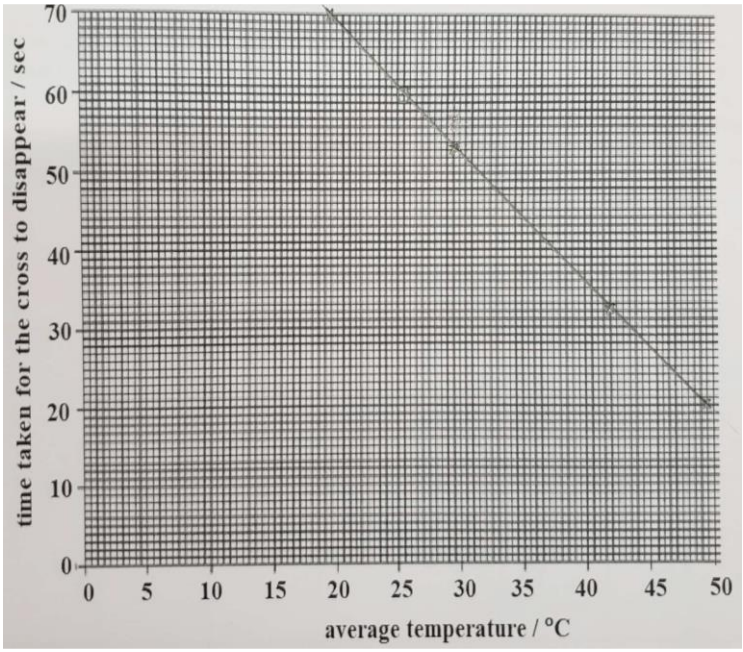
d. Identify the experiment at which the rate of reaction was lowest. Explain your answer in terms of kinetic molecular theory.

e. A student added hydrochloric acid to sodium thiosulfate solution using a pipette instead of the measuring cylinder.
Would you recommend this step? Explain.

f. The use of pipettes or burettes to measure volumes of the solutions used in these experiments would improve the accuracy of the results.
Suggest another way to improve the accuracy of the results of these experiments. Explain your answer.

a. Avg temperature: 19, 29, 44, 49 for each experiment
Time: 70, 53, 32, 20

b.



c. 26°C (from the graph)

d. Experiment 1 , experiment run at the lowest temperature ⇒ reacting particles have least (kinetic) energy ⇒ move slower ⇒ less frequent collisions

e. Not recommended addition using a pipette would be slower affecting time of reaction

f. Insulation or use a lid to reduce heat losses

CQD32. [T] Rate of chemical reaction is affected by changing the surface area of reactants.

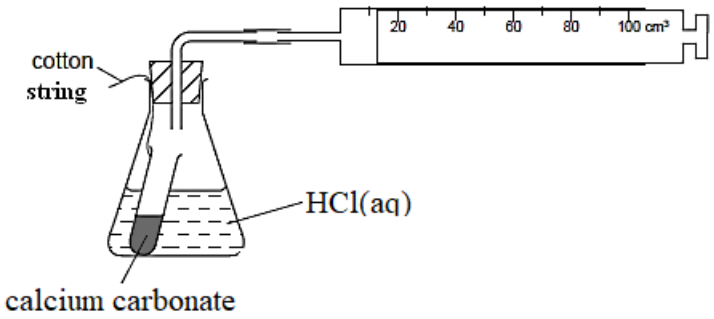
Calcium carbonate reacts with dilute hydrochloric acid to produce carbon dioxide. The rate of the reaction can be followed by measuring the volume of carbon dioxide released.

Design an experiment to investigate the effect of surface area on the rate of this reaction by measuring the volume of carbon dioxide released.

You are provided with lumps of calcium carbonate, powdered calcium carbonate, dilute hydrochloric acid and common laboratory apparatus.

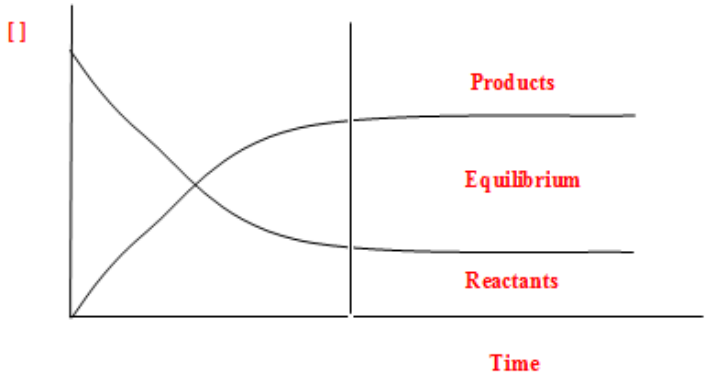
Draw a labelled diagram of the apparatus you would use.

• Weighed amount / x gram of lumps calcium carbonate
• Add known volume of dilute hydrochloric acid
• gas syringe /
• Use of stop-clock / timer
• Measure volume of carbon dioxide at fixed time or time for a fixed volume to be made
• Repeat using powdered calcium carbonate
• Compare results / conclusion



Chapter 1-Chemical Equilibrium

1. As any reaction proceeds, [reactants] decreases and [products] increases.



2. **Reversible reaction:** a reaction which can go in opposite direction depending on the conditions. It is indicated by a reversible arrow: ⇄.
Typical reversible reactions are hydration of salts, when heated the forward reaction occurs, when water is added, the reverse reaction occurs:

$CuSO_4.5H_2O(s) \underset{water}{\overset{heat}{\rightleftarrows}} CuSO_4(s) + 5H_2O(l)$

hydrated copper(II) sulfate — ***anhydrous*** copper(II) sulfate

$CoCl_2.6H_2O(s) \underset{water}{\overset{heat}{\rightleftarrows}} CoCl_2(s) + 6H_2O(l)$

pink solid — blue solid

hydrated **cobalt(II) chloride** — ***anhydrous*** **cobalt(II) chloride**

As a matter of fact, the above two reactions can be used to test for the presence of water. Note that this test can indicate water's presence but NOT its purity.

Hydrated and anhydrous copper(II) sulfate

Hydrated and anhydrous cobalt(II) chloride

From www.yooniqimages.com)
https://www.sciencephoto.com/media/1149592/view/anhydrous-and-hydrated-cobalt-chloride

3. Water of crystallization is defined as the water molecules present in hydrated crystals such as $CuSO_4.5H_2O(s)$ and $CoCl_2.6H_2O(s)$.

4. A hydrated substance is one that is chemically combined with water while an anhydrous substance is one containing no water.

5. **Equilibrium is reached** when no more changes can be observed. It is **recognized** by constancy of macroscopic properties in a closed system at constant temperature.

6. Macroscopic properties are observable properties or measurable properties like pressure, concentration, color, size, volume and mass.

7. A steady state describes a system that has some constant properties but equilibrium does not exist.

CQD 1. Which of the following systems constitute steady state situations, and which are at equilibrium? For each, a constant property is indicated.

a. An unopened bottle of sparkling water (a bottle containing CO_2 and water).
b. A Bunsen lit with a constant blue flame color.

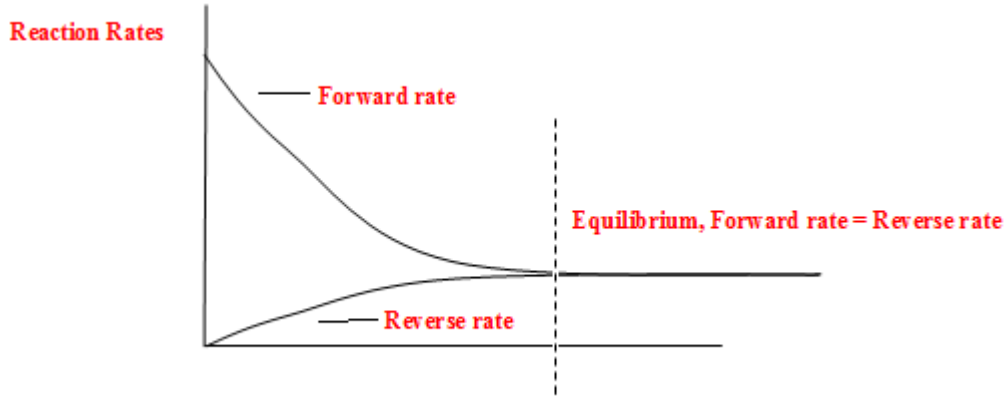
8. **Closed system** is a system that does not allow exchange of matter but allows exchange of energy.

9. A reaction is at **equilibrium** when:

- the system is closed
- the [reactants] and [products] are constant (**NOT** necessarily equal)
- The rate of the forward reaction equals the rate of the backward reaction

CQD 2. [G] What happens when a system reaches equilibrium?

10. **At equilibrium**:



11. Equilibrium is dynamic in nature since at equilibrium two microscopic processes are occurring in opposite direction at the same rate resulting in no observable macroscopic changes. In other words, reactants are continuously changing into products and products are continuously changing back into reactants.

12. Equilibrium may not be reached in certain equilibrium system because the activation energy of the forward reaction is too high.

13. Factors that affect the position of equilibrium include;

- Temperature
- Concentration
- Pressure / Volume for gaseous systems

14. The position of equilibrium indicates how far a reaction has gone to favour reactants or products. When the position of equilibrium shifts to the right, then products are favoured and when the position of equilibrium shifts to the left, then reactants are favoured.

15. Le Chatelier's Principle (LCP): if a system at equilibrium is subjected to a change, processes will occur to partially counteract the imposed change.

- If the concentration of a reactant is increased, equilibrium shifts to the right.
- If the concentration of a product is increased, equilibrium shifts to the left.
- Changing the amount of a pure liquid or a pure solid will not alter the position of equilibrium.
- In a gaseous system, increasing the volume is the same as decreasing pressure which shifts the equilibrium to the side with more **gaseous** moles.
- In a gaseous system, decreasing the volume is the same as increasing pressure which shifts the equilibrium to the side with less **gaseous** moles.
- Increasing the temperature of an exothermic reaction will shift the reaction to the left.
- Decreasing temperature of an exothermic reaction will shift the reaction to the right.
- Increasing the temperature of an endothermic reaction will shift the reaction to the right.
- Decreasing temperature of an endothermic reaction will shift the reaction to the left.
- Catalysts DO NOT affect the position of equilibrium. Catalysts only allow the reaction to achieve equilibrium faster.

CQD 3.

a. Predict what would happen to equilibrium concentrations of $H_2O(g)$ and $H_2(g)$ respectively if the partial pressure of O_2 is increased by injecting more $O_2(g)$ into the reaction chamber at 273°C.

$$H_2O(g) \rightleftarrows H_2(g) + \tfrac{1}{2}O_2(g) \qquad \Delta H = +242\ kJ$$

b. $2SO_3(g) \leftrightarrows 2SO_2(g) + O_2(g)$

After the equilibrium represented above is established, some pure $SO_2(g)$ is injected into the reaction vessel at constant temperature. Which of the following has a lower value compared to its value at the original equilibrium?

A. The total pressure in the reaction vessel
B. The amount of $SO_3(g)$ in the reaction vessel
C. The amount of $O_2(g)$ in the reaction vessel
D. The amount of $SO_2(g)$ in the reaction vessel

c. The chromate and dichromate ions are set up in equilibrium as follows:

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$
yellow orange

What happens in each of the following cases? Complete the table below

Change	Equilibrium shifts to	Color change	Change to $[Cr_2O_7^{2-}]$
Increase $[H^+]$			
Increase in $[CrO_4^{2-}]$			
Decrease $[H^+]$			
Add NaOH			
Add HCl			

d. Predict what will happen to the equilibrium concentration of CO_2 (g) if some $CaCO_3$ (s) is added to the system:

$$CaCO_3(s) \leftrightarrows CaO(s) + CO_2(g)$$

e. Predict what would happen to equilibrium concentrations of $H_2O(g)$ and $H_2(g)$ respectively if the pressure in the reaction chamber at 273°C is increased by adding (unreactive) He(g).

$$H_2O(g) \rightleftarrows H_2(g) + \tfrac{1}{2} O_2(g) \qquad \Delta H = +242 \text{ kJ}$$

f. The equilibrium system represented below is contained in a sealed, rigid vessel.

$$4HCl(g) + O_2(g) \rightleftarrows 2H_2O(g) + 2Cl_2(g) \quad \Delta H^o = -113 \text{ kJ.mol}^{-1}$$

Which of the following will decrease if the temperature of the mixture is raised?

- $[HCl\ (g)]$
- $[O_2\ (g)]$
- $[Cl_2\ (g)]$
- $[HCl\ (g)]$ and $[Cl_2\ (g)]$
- $[O_2\ (g)]$, $[HCl\ (g)]$ and $[Cl_2\ (g)]$

g. Consider the following system: $2SO_2(g) + O_2(g) \leftrightarrows 2SO_3(g)$
State the effect of adding a catalyst to the system on the position of equilibrium?

CQD 5. [G] Consider the following equilibrium system.

$$Fe(aq) + SCN(aq)^{3+} \rightleftarrows FeSCN^{2+}(aq)$$

Fill in the table to predict the effect of the given changes on the equilibrium and the amount of $FeSCN^{2+}$ produced.

Imposed change	Equilibrium shifts	$[FeSCN^{2+}]$
Removal of Fe^{3+}		
Addition of SCN^-		

CQD 6. [G] Fill in the table to determine the effect of increasing pressure on the given equilibrium systems and the amount of product in each.

Equilibrium system	Equilibrium	Amount of product(s)
$H_2(g) + Cl_2(g) \rightleftarrows 2HCl(g)$		
$2H_2(g) + O_2(g) \rightleftarrows 2H_2O(g)$		
$2NH_3(g) \rightleftarrows 3H_2(g) + N_2(g)$		

CQD 7. [G] Use Le Chatelier's Principle to predict the effect of an increase in temperature on the given equilibrium systems and the equilibrium concentrations of the products.

Equilibrium system	Equilibrium shifts	[product]
$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g) + 192\ kJ$		
$N_2(g) + O_2(g) + 180.5\ kJ \rightleftarrows 2NO(g)$		
$CO(g) + H_2O(g) \rightleftarrows CO_2(g) + H_2(g) \quad \Delta H = -27.6\ kJ$		

CQD 8. [G] Use Le Chatelier's Principle to predict the effect of the addition of a catalyst on the given equilibrium system and the equilibrium concentrations of the products.

$$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g) + 192\ kJ$$

CQD 9. [G] Use Le Chatelier's Principle to predict the effect of the addition of an inert gas on the given equilibrium system taking place in a rigid sealed container and the equilibrium concentrations of the products.

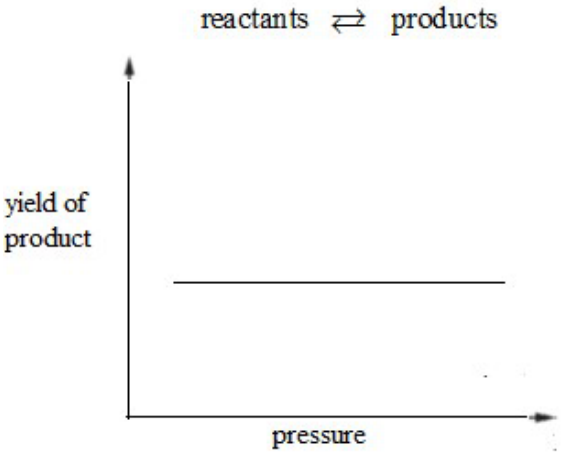
$$N_2(g) + O_2(g) + 180.5\ kJ \rightleftarrows 2NO(g)$$

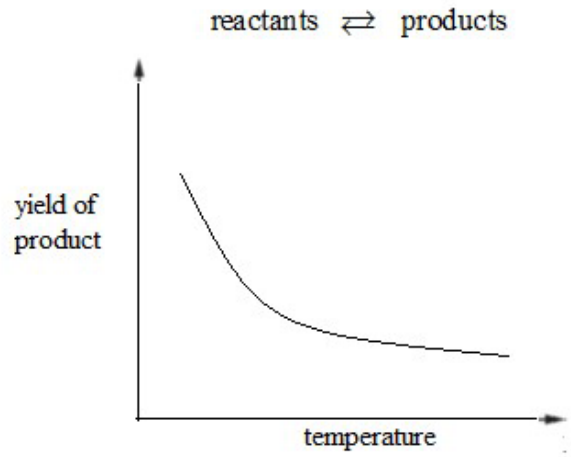
CQD 4. The "yield" can be defined as the amount of product formed by a reaction. How can we increase the yield of ammonia production?

Ammonia is produced by combining nitrogen and hydrogen gases as follows:

$$N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g) \quad \Delta H = -92 \text{ kJ}$$

CQD 14. [T] The graphs below show how the yield of product changes as pressure and temperature varies.





Which of the following reactions is represented in the graphs?

A. $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$ $\Delta H < 0$
B. $2SO_3(g) \rightleftarrows 2SO_2(g) + O_2(g)$ $\Delta H > 0$
C. $H_2CO_3(aq) \rightleftarrows H_2O(l) + CO_2(g)$ $\Delta H > 0$
D. $N_2(g) + O_2(g) \rightleftarrows 2NO(g)$ $\Delta H < 0$
E. $2H_2O(l) \rightleftarrows O_2(g) + 2H_2(g)$ $\Delta H > 0$

CQD 15. [T] Carbon dioxide and hydrogen gases react at 75 atm and 250°C to produce methyl alcohol industrially.

$CO(g) + 2H_2(g) \rightleftarrows CH_3OH(g)$, the backward reaction is endothermic.

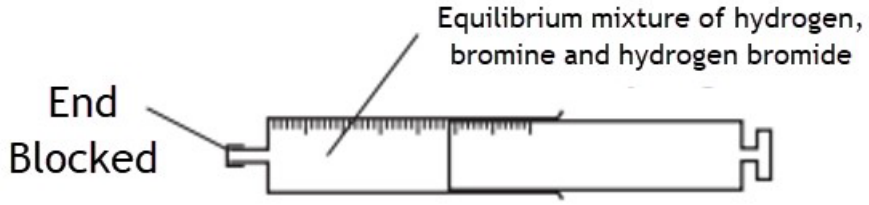
a. Write a source of hydrogen for this industrial process.
b. Complete the table using the words decreases, increases, or no change.

	Effect on the rate of the forward reaction	Effect on the equilibrium yield of $CH_3OH(g)$
Increasing the pressure		
Decreasing temperature		
Adding a catalyst		

CQD 16. [T] Bromine and hydrogen gases react together in a reversible reaction. Hydrogen bromide is formed.

$$\underset{\text{Colorless}}{H_2(g)} + \underset{\text{Brown}}{Br_2(g)} \rightarrow \underset{\text{Colorless}}{2HBr(g)}$$

The forward reaction is exothermic.
A gas syringe containing an equilibrium mixture of hydrogen, bromine and hydrogen bromide gases was sealed and heated to 300°C. The equilibrium mixture was a pale brown color.



a. What is meant by the term equilibrium?
b. As the syringe plunger was pressed while keeping the other end blocked, the pressure increased. It was shown that the position of the equilibrium did not change. The color of the gaseous mixture turned darker brown.

i. Give a reason why the position of the equilibrium did not change.
ii. Suggest why the gaseous mixture turned darker brown, even though the position of the equilibrium did not change.

c. The temperature of the gas syringe was increased to 400°C.

i. What happened to the position of the equilibrium when the temperature of the gas syringe was increased from 250°C to 400°C?
ii. What happened to the rate of the forward reaction and the rate of the backward reaction when the temperature of the gas syringe was increased from to 250°C to 400°C?

- Rate of the forward reaction
- Rate of the backward reaction

16. **Mass Action Expression**: For a general reaction: $\alpha A + \beta B \leftrightarrows \Upsilon C + \Delta D$
    Mass action expression = Q = $\frac{[C]^{\Upsilon}[D]^{\Delta}}{[A]^{\alpha}[B]^{\beta}}$
17. **At equilibrium**: mass action expression is constant and is given a special name, equilibrium constant, K_{eq}

$$K_{eq} = \frac{[C]^{\Upsilon}[D]^{\Delta}}{[A]^{\alpha}[B]^{\beta}}$$

18. **Concentrations of solids and liquids are NOT included in the equilibrium expression**. These values are constant and are incorporate in the value of K_{eq} directly.
19. Temperature and the nature of solvent are the only values which determine the value of the equilibrium constant.

Exercise:
Write the equilibrium constant expression of the following systems:

A. $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$
B. $2SO_3(g) \rightleftarrows 2SO_2(g) + O_2(g)$
C. $H_2CO_3(aq) \rightleftarrows 2H^+(aq) + CO_3^{2-}(aq)$
D. $N_2(g) + O_2(g) \rightleftarrows 2NO(g)$
E. $2H_2O(g) \rightleftarrows O_2(g) + 2H_2(g)$

Chapter 2-Solubility

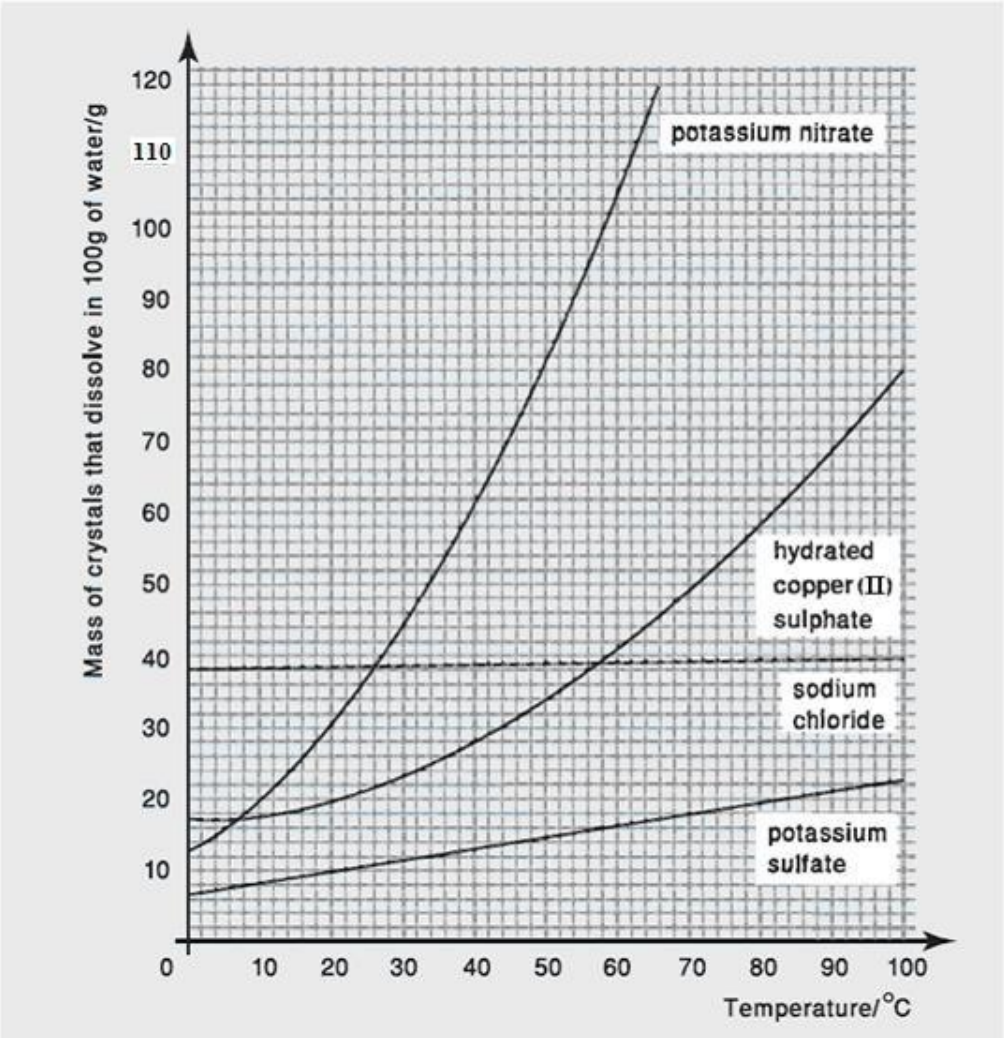
1. **Molecular compounds**: those made up of molecules, for eg. glucose.
2. **Most** molecular compounds remain molecular when dissolved in water (ex glucose).
3. **Some** molecular compounds hydrolyze into ions when dissolved in water, for eg. hydrogen halides, H_2SO_4, HNO_3.
4. **Ionic compounds**: form when metals and nonmetals react.
5. **Ionic compounds:** dissociate into ions in aqueous solutions.

6. **Solubility:** is the molar concentration of a saturated solution. It is also defined in terms of max mass of solute that can dissolve in 100g solvent.
7. **Saturated solution**: The solvent contains as much solute as it can hold at equilibrium at a given temperature. The solid must be in contact with the liquid.
8. If a saturated solution at a certain temperature is cooled, the solid precipitates out.
9. An insoluble solid is always in equilibrium with its ions in solution.
10. In general heat increase solubility of solids and decrease solubility of gases.

CQD 1. Propose two methods for increasing the rate at which salt dissolves in water.

11. Use a solubility curve to:
    - read the solubilities of solids at different temp.
    - determine the increase in solubility when heated
    - determine the mass of solid formed when a saturated solution is cooled
    - determine mass of water needed to dissolve a specific mass of solid at a specific temp
    - determine mass of solid that will dissolve in a specific mass of water at a specific temp

CQD 2. Given:



a. What is the solubility of potassium nitrate at 20ºC? at 50ºC?
b. By how much will the solubility of potassium nitrate increase, if the temperature is increased from 20°C to 50°C?
c. What mass of potassium nitrate forms, if a saturated solution in 100 g of water is cooled from 40ºC to 10ºC? from 60ºC to 20ºC?
d. What mass of water is needed to dissolve 50 g of potassium nitrate at 40ºC?
e. What mass of potassium sulfate will dissolve in 2 kg of water at 20ºC?

CQD 3. Choose the correct answer and complete the following equation in each.

a) Glucose is (molecular / ionic) and it dissolves in water as such: $C_6H_{12}O_6(s) \rightarrow$
b) NaCl is (molecular / ionic) and it dissolves in water as such: $NaCl(s) \rightarrow$
c) HCl gas is (molecular / ionic) and it dissolves in water as such: $HCl(g) \rightarrow$

12. **Electrolytes** are substances that dissolve in water to form conducting solutions.
13. **Cations** are positively charged ions, **anions** are negatively charged ions.
14. SUMMARY **of rules to remember:**

- Sodium, potassium, ammonium salts are soluble
- Nitrates and acetates are soluble
- Halides are soluble except lead and silver
- Sulfates are soluble except barium, calcium and lead
- Carbonates are insoluble, except sodium, potassium and ammonium
- Hydroxides are insoluble, except sodium, potassium, ammonium and calcium (partially)
- $PbCl_2$ is **soluble** in hot water and insoluble in cold.

CQD 7. [G] Classify the following salts as soluble or slightly soluble:
a. all salts containing nitrates NO_3^-.
b. all salts of alkali metal ions.
c. group 2 metal carbonates.
d. chlorides of silver and lead.

15. Use the rules to predict if a precipitate will form upon mixing two given solutions.

CQD 4. If we mix the following solutions:

- state what will be observed
- name of precipitate
- write a net ionic equation of the reaction occurring

a. $Ba(NO_3)_2$ and Na_2SO_4
b. $AgNO_3$ and NaCl
c. $Ba(NO_3)_2$ and NaCl
d. $Ca(NO_3)_2$ and Na_2CO_3

16. Colors of common compounds you need to remember:

- ✓ All group 1 salts are white and their solutions are colourless
- ✓ All carbonate salts are white except copper(II) carbonate which is green
- ✓ All sulfide salts are white except CuS, FeS and PbS which are black
- ✓ Iron (II) hydroxide, $Fe(OH)_2$, is green
- ✓ Iron (III) hydroxide, $Fe(OH)_3$, is red- brown
- ✓ Chromium (III) hydroxide, $Cr(OH)_3$, is green
- ✓ Barium sulfate is white
- ✓ Silver chloride is white , silver bromide is creamy white and silver iodide is yellow
- ✓ Lead (II) chloride is white, lead (II) bromide is creamy, lead(II) iodide is yellow
- ✓ Copper(II) hydroxide is blue
- ✓ Aluminium hydroxide is gelatinous white
- ✓ Zinc hydroxide is white
- ✓ Calcium hydroxide is white

CQD 5. A solution is known to contain $Pb(NO_3)_2$ and $Mg(NO_3)_2$. It is required to remove the lead(II) ions from the solution without removing magnesium ions. This can be done by adding just the right amount of

a. NaCl.
b. HNO_3.
c. NH_3.

CQD 6 What would be the weight of the precipitate formed when 100. mL of 0.500 M NaCl is added to 50.0 mL of 0.100 M $AgNO_3$?

17. Demonstration: flame tests

18. <u>Demonstration</u>: test for halide ions using silver nitrate solution followed with ammonia

19. <u>Demonstration</u>: test for NH_4^+ using NaOH and heat

20. Tests: the tables below will be provided in paper 6 of IGCSE exam.

Tests for anions

anion	test	test result
carbonate, CO_3^{2-}	add dilute acid, then test for carbon dioxide gas	effervescence, carbon dioxide produced
chloride, Cl^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
bromide, Br^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	cream ppt.
iodide, I^- [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate, NO_3^- [in solution]	add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
sulfate, SO_4^{2-} [in solution]	acidify with dilute nitric acid, then add aqueous barium nitrate	white ppt.
sulfite, SO_3^{2-}	add a small volume of acidified aqueous potassium manganate(VII)	the acidified aqueous potassium manganate(VII) changes colour from purple to colourless

Tests for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium, Al^{3+}	white ppt., soluble in excess, giving a colourless solution	white ppt., insoluble in excess
ammonium, NH_4^+	ammonia produced on warming	-
calcium, Ca^{2+}	white ppt., insoluble in excess	no ppt. or very slight white ppt.
chromium(III), Cr^{3+}	green ppt., soluble in excess	green ppt., insoluble in excess
copper(II), Cu^{2+}	light blue ppt., insoluble in excess	light blue ppt., soluble in excess, giving a dark blue solution
iron(II), Fe^{2+}	green ppt., insoluble in excess, ppt. turns brown near surface on standing	green ppt., insoluble in excess, ppt. turns brown near surface on standing
iron(III), Fe^{3+}	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc, Zn^{2+}	white ppt., soluble in excess, giving a colourless solution	white ppt., soluble in excess, giving a colourless solution

Tests for gases

gas	test and test result
ammonia, NH_3	turns damp red litmus paper blue
carbon dioxide, CO_2	turns limewater milky
chlorine, Cl_2	bleaches damp litmus paper
hydrogen, H_2	‘pops’ with a lighted splint
oxygen, O_2	relights a glowing splint
sulfur dioxide, SO_2	turns acidified aqueous potassium manganate(VII) from purple to colourless

Flame tests for metal ions

metal ion	flame colour
Lithium, Li^+	red
sodium, Na^+	yellow
potassium, K^+	lilac
calcium, Ca^{2+}	orange-red
barium, Ba^{2+}	light green
copper(II), Cu^{2+}	blue-green

CQD 8 [G] Given four separate test tubes labelled A, B, C, and D containing solutions of four different potassium halides. A few milliliters of an aqueous solution of silver nitrate were added to each of the test tubes and the following observations were recorded.

Test tube	**Observation upon addition of few milliliters of $AgNO_3$(aq)**
A	Formation of yellow precipitate
B	Formation of white precipitate that turns to violet then to black when exposed to light.
C	Formation of creamy precipitate that darkens when exposed to light.
D	No change

Give the chemical formula for the halide present in each of the given test tubes.

CQD 9 [G] Fill in the table to identify the ion present in each solution after conducting a flame test.

Solution	**Flame Color**	**Cation**
A	**Yellow**	
B	**Lilac/ Violet**	
C	**Crimson red**	
D	**Blue-green**	
E	**Pale green**	
F	**Brick Red**	

CQD 10 [G] Seven different solutions each containing one of the cations: Cr^{3+}, Ca^{2+}, Zn^{2+}, Cu^{2+}, Fe^{2+}, Al^{3+}, or Fe^{3+} were tested using sodium hydroxide or ammonia. The observations were recorded in the table below. Match these ions to the below observations.

Observation with solution tested	**Cation present in solution**
Reddish-brown precipitate with sodium hydroxide, does not dissolve in excess base.	
Dirty-green precipitate with sodium hydroxide, does not dissolve in excess base.	
Blue precipitate with sodium hydroxide, does not dissolve in excess sodium hydroxide.	
White precipitate with sodium hydroxide, and no precipitate forms with ammonia.	
White precipitate with ammonia, does not re-dissolve in excess ammonia.	
White precipitate with sodium hydroxide, re-dissolve in excess ammonia.	
Green precipitate with sodium hydroxide, soluble in excess sodium hydroxide.	

CQD 11 [G] Given the following cations. Ag^{+}, Ca^{2+}, Na^{+}, Zn^{2+}, Ba^{2+}.

a. Which of the given ions form a white precipitate upon addition of an aqueous solution of potassium sulfate, K_2SO_4?
b. Which of the given ions forms a white precipitate upon addition of an aqueous solution of sodium chloride, NaCl?
c. Which of the given ions forms a creamy precipitate upon addition of an aqueous solution of potassium bromide, KBr?

CQD 12 [G] Given the following ions: SO_3^{2-}, SO_4^{2-}, CO_3^{2-}, NH_4^{+}, Al^{3+}, NO_3^{-}.

a. Which of the given ions releases a gas that turns limewater milky upon reacting with an acid?
b. Which of the given ions releases a pungent smell gas that turns moist red litmus blue upon reacting with a base?
c. Which of the given ions releases a gas that turns moist red litmus to blue when heated with aluminum powder and a base?
d. Which of the given ions releases a gas with a chocking odor upon addition of an acid?

CQD 16 [T] Many insoluble salts can be prepared by precipitation reaction between soluble salts. All potassium salts and all nitrate salts are soluble.

a. A student adds a potassium sulfate, K_2SO_4, solution to a barium nitrate, $Ba(NO_3)_2$, solution. A white precipitate is formed.

i. What is the chemical formula of the white precipitate?
ii. Write the chemical equation for the reaction between potassium sulfate and barium nitrate.

b. Another student adds potassium sulfate, K_2CO_3, solution to a silver nitrate, $AgNO_3$, solution. A white precipitate is also formed.

i. What is the chemical formula of the white precipitate?
ii. Write the net ionic equation of the reaction between potassium carbonate and silver nitrate.

c. Barium nitrate solution produces a white precipitate with both carbonate ions and sulfate ions. Why acidified barium nitrate solution is used to distinguish between carbonate ions and sulfate ions?

November 2020/P62

3 Solid **Q** and solid **R** were analysed. Solid **Q** was zinc carbonate.
Tests were done on each solid.

tests on solid Q

Complete the expected observations.

(a) Solid **Q** was placed in a boiling tube. About 10 cm^3 of dilute sulfuric acid was added to the boiling tube. Any gas produced was tested.
The contents of the boiling tube were kept for **(c)**.

observations .....

.....

..... [3]

(b) Identify the gas given off in **(a)**.

..... [1]

(c) The reaction mixture from **(a)** was filtered.
The filtrate was solution **S**. 1 cm depth of solution **S** was poured into a boiling tube.

(i) Aqueous sodium hydroxide was added dropwise and then in excess to solution **S** in the boiling tube.

observations .....

.....

..... [2]

(ii) Explain why it is **not** possible to identify the cation contained in solution **S** from your observations in **(c)(i)**.

.....

..... [1]

(iii) Suggest an additional test that can be done on solution **S** to confirm the cation was Zn^{2+}.

.....

..... [1]

tests on solid R

Tests were done and the following observations were made.

tests on solid **R**	observations
test 1 A flame test was done on solid **R**.	yellow flame
Solid **R** was dissolved in distilled water to produce solution **R**. The solution was divided into two equal portions in two test-tubes. **test 2** About 1 cm^3 of dilute nitric acid followed by a few drops of aqueous silver nitrate were added to the first portion of solution **R**.	yellow precipitate formed
test 3 The second portion of solution **R** was added to 1 cm^3 of aqueous bromine in a test-tube.	the solution changed colour from orange to brown

(d) Identify solid **R**.

.....

..... [2]

[Total: 10]

November 2022/P62

3 Solid **I** and solid **J** were analysed.
Tests were done on each substance.

tests on solid I

tests	observations
test 1 Dilute hydrochloric acid was added to a boiling tube containing solid **I**. Any gas produced was tested.	effervescence was seen, the solid dissolved to form a colourless solution the gas turned limewater milky
test 2 A flame test was carried out on the solution formed in **test 1**.	a red flame was seen

(a) Identify the gas made in **test 1**.

..... [1]

(b) Identify solid **I**.

.....

..... [2]

tests on solid J

Solid **J** was aluminium chloride.

Solid **J** was dissolved in water to form solution **J**. Solution **J** was divided into four approximately equal portions in four test-tubes.

(c) Aqueous sodium hydroxide was added dropwise and then in excess to the first portion of solution **J**.

observations .....

..... [2]

(d) Aqueous ammonia was added dropwise and then in excess to the second portion of solution **J**.

observations .....

..... [2]

(e) About 1 cm depth of dilute nitric acid followed by a few drops of aqueous barium nitrate were added to the third portion of solution **J**.

observations ..... [1]

(f) About 1 cm depth of dilute nitric acid followed by a few drops of aqueous silver nitrate were added to the fourth portion of solution **J**.

observations ..... [1]

[Total: 9]

March 2021/P62

3 Two solids, solid **C** and solid **D**, were analysed.
Tests were done on each solid.

tests on solid C

Tests were carried out and the following observations were made.

tests	observations
test 1 A flame test was carried out on solid **C**.	a red flame was seen
Solid **C** was dissolved in distilled water to produce solution **C**. **test 2** About 5 cm^3 of aqueous sodium hydroxide was added to solution **C**.	no change
test 3 A piece of aluminium foil was added to the mixture formed in **test 2**.The mixture was warmed gently and any gas produced was tested.	effervescence was seen; damp red litmus paper turned blue

(a) Name the gas that turned the damp red litmus paper blue in **test 3**.

..... [1]

(b) Identify solid **C**.

.....

..... [2]

tests on solid D

Solid **D** was aluminium sulfate.

Complete the expected observations.

Solid **D** was dissolved in water to form solution **D**. Solution **D** was divided into four approximately equal portions in four test-tubes.

(c) Aqueous sodium hydroxide was added dropwise and then in excess to the first portion of solution **D**.

observations .....

..... [2]

(d) Aqueous ammonia was added dropwise and then in excess to the second portion of solution **D**.

observations .....

..... [2]

(e) About 1 cm^3 of dilute nitric acid and a few drops of aqueous silver nitrate were added to the third portion of solution **D**.

observations ..... [1]

(f) About 1 cm^3 of dilute nitric acid and a few drops of aqueous barium nitrate were added to the fourth portion of solution **D**.

observations ..... [1]

[Total: 9]

June 2023/P62

3 A student tests two substances: solution **F** and solid **G**.

Tests on solution F

Table 3.1 shows the tests and the student's observations for solution **F**. The student divides solution **F** into three portions.

Table 3.1

tests	observations
test 1 Do a flame test on the first portion of solution **F**.	light green colour
test 2 To the second portion of solution **F**, add a 1 cm depth of aqueous sodium hydroxide and a piece of aluminium foil. Warm the mixture gently and test any gas produced.	effervescence was seen; the gas turned damp red litmus paper blue
test 3 To the third portion of solution **F**, add a 1 cm depth of dilute nitric acid followed by a few drops of aqueous silver nitrate.	no change

(a) Describe how to do the flame test used in **test 1**.

.....

.....

..... [2]

(b) Identify the gas given off in **test 2**.

..... [1]

(c) Identify solution **F**.

.....

..... [2]

(d) State what would be observed if the student adds dilute sulfuric acid to another portion of solution **F**.

observations .....

..... [1]

Tests on solid G

Solid **G** is iron(II) carbonate.

(e) About 10 cm^3 of dilute sulfuric acid is added to solid **G**.
Any gas given off is tested.

observations .....

.....

..... [2]

(f) To the product from **(e)**, aqueous sodium hydroxide is added dropwise until in excess.

observations adding dropwise .....

observations in excess .....
[2]

[Total: 10]

June 2023/P61

3 A student tests two solids: solid **E** and solid **F**.

Tests on solid E

Table 3.1 shows the tests and the student's observations.

Table 3.1

tests	observations
test 1 Gently heat half of solid **E** in a boiling tube.	a solution forms, steam is given off and condensation forms at the top of the tube
test 2 Dissolve the remaining solid **E** in water to form solution **E**. Divide solution **E** into three portions. To the first portion of solution **E**, add aqueous sodium hydroxide dropwise and then in excess.	a brown precipitate forms which remains when excess is added
test 3 Warm the product of **test 2** and test any gas produced.	the gas turns red litmus paper blue
test 4 To the second portion of solution **E**, add 1 cm depth of dilute nitric acid followed by a few drops of aqueous silver nitrate.	no change
test 5 To the third portion of solution **E**, add 1 cm depth of dilute nitric acid followed by a few drops of aqueous barium nitrate.	white precipitate

(a) State what conclusion can be made about solid **E** from the observations in **test 1**.

..... [1]

(b) Identify the gas produced in **test 3**.

..... [1]

(c) State what conclusion can be made about solid **E** from the observations in **test 4**.

..... [1]

(d) Identify the **three** ions in solid **E**.

.....

..... [3]

Tests on solid F

Solid **F** is zinc sulfite.

Complete the expected observations.

The student dissolves solid **F** in water to form solution **F**.

The student divides solution **F** into three portions.

(e) To the first portion of solution **F**, the student adds aqueous ammonia dropwise until it is in excess.

observations adding dropwise .....

observations in excess .....

[2]

(f) To the second portion of solution **F**, the student adds a few drops of acidified aqueous potassium manganate(VII).

observations .....

..... [1]

(g) To the third portion of solution **F**, the student adds 1 cm depth of dilute nitric acid followed by a few drops of aqueous barium nitrate.

observations .....

..... [1]

[Total: 10]

November 2022/P61

3 Two substances, solid **W** and solid **X**, were analysed. Solid **W** was zinc bromide.

tests on solid W

Complete the expected observations.

Solid **W** was dissolved in water to form solution **W**. Solution **W** was divided into three equal portions.

(a) To the first portion of solution **W**, aqueous ammonia was added dropwise and then in excess.

observations .....

..... [2]

(b) To the second portion of solution **W**, 1 cm^3 of dilute nitric acid followed by a few drops of aqueous barium nitrate were added.

observations ..... [1]

(c) To the third portion of solution **W**, 1 cm^3 of dilute nitric acid followed by a few drops of aqueous silver nitrate were added.

observations ..... [1]

tests on solid X

tests	observations
test 1 About 1 g of solid **X** was placed in a boiling tube and heated strongly. A strip of filter paper soaked in acidified aqueous potassium manganate(VII) solution was held at the mouth of the boiling tube.	the acidified aqueous potassium manganate(VII) turned from purple to colourless
The remaining solid **X** was dissolved in water to form solution **X**. Solution **X** was divided into three equal portions. **test 2** 1 cm^3 of dilute nitric acid and a few drops of aqueous silver nitrate were added to the first portion of solution **X**.	no change
test 3 1 cm^3 of dilute nitric acid and a few drops of aqueous barium nitrate were added to the second portion of solution **X**.	a white precipitate formed
test 4 Aqueous sodium hydroxide was added dropwise and then in excess to the third portion of solution **X**.	a green precipitate formed and remained in excess

(d) (i) Name the gas given off in **test 1**.

..... [1]

(ii) Water vapour is also given off in **test 1**.

Give a chemical test for water and the expected observation if water is present.

substance used .....

observation .....

[2]

(e) Identify solid **X**.

.....

..... [2]

[Total: 9]

June 2021/P62

3 Solution **G** and solid **H** were analysed.

tests on solution G

tests	observations
Solution **G** was divided into three equal portions in three test-tubes. **test 1** Sodium hydroxide was added dropwise and then in excess to the first portion of solution **G**.	white precipitate which did not dissolve in excess
test 2 About 1 cm^3 of dilute nitric acid followed by a few drops of aqueous silver nitrate were added to the second portion of solution **G**.	yellow precipitate
test 3 About 10 cm^3 of aqueous hydrogen peroxide was added to the third portion of solution **G**. The gas produced was tested.	the mixture became brown and bubbled; the gas relit a glowing splint

(a) Identify the gas produced in **test 3**.

..... [1]

(b) Use the results of **test 1** and **test 2** to identify solution **G**.

.....

.....

..... [2]

tests on solid H

Solid **H** was hydrated copper(II) sulfate.

Complete the expected observations.

(c) About half of solid **H** was placed in a boiling tube and heated using a Bunsen burner.

observations .....

..... [2]

(d) A flame test was carried out on solid **H**.

observations ..... [1]

The remaining solid **H** was placed in a boiling tube. About 10 cm^3 of distilled water was added to the boiling tube. The tube was shaken to dissolve solid **H** and form solution **H**.

Solution **H** was divided into two approximately equal portions in two test-tubes.

(e) Aqueous ammonia was added dropwise and then in excess to the first portion of solution **H**.

observations .....

.....

..... [3]

(f) Approximately 1 cm^3 of dilute nitric acid followed by a few drops of aqueous barium nitrate were added to the second portion of solution **H**.

observations ..... [1]

[Total: 10]

Chapter 3-Acids, Bases and Salts

1. **An electrolyte:** is a solution that dissolves in water to form conducting solutions. The solutions conduct because they contain ions.

(Substance that conducts electricity when molten or dissolved in water because of the presence of delocalized ions/ substance which undergoes electrolysis)

2. **A strong electrolyte:** is one that dissolves and exclusively dissociates in aqueous solutions into ions. (eg. NaCl, HCl, NaOH)

3. **A weak electrolyte:** is one that partially dissociates in aqueous solutions into ions (ex CH_3COOH)

4. Water is a weak electrolyte, with an equilibrium constant, $K_w = 10^{-14}$. At 25°C, in pure water $[OH^-] = [H^+] = 1.00 \times 10^{-7}$ M

5. According to the **Arrhenius definition,** an acid is a substance that increases the concentration of H^+ ions in water. A base is a substance that that increases the concentration of OH^- ions in water.
6. According to the **Bronsted Lowry definition**, an acid is a proton donor and a base is a proton acceptor.
7. Many acids are dangerous as they are corrosive, that is they can "eat away" at other material.
8. A strong acid is one that dissociates completely while a weak acid is one that dissociates partially.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$
$$CH_3COOH(aq) \rightleftarrows CH_3COO^-(aq) + H^+(aq)$$

9. Similarly, a strong base is one that dissociates completely while a weak base is one that dissociates partially.
10. Examples of strong acids are hydrochloric acid, hydrobromic acid, hydroiodic acid, sulfuric acid, nitric acid.

11. $[H^+]$ and $[OH^-]$ can be easily calculated in a given strong acid/strong base solution of known conc.

CQD 1 [G] Calculate the $[H^+]$ and the $[OH^-]$ in a 0.010 M solution of LiOH at 25°C.

CQD 9 [G]

a. Calculate the $[H^+]$ and the $[OH^-]$ in a 0.010 M solution of HCl.

b. Calculate the $[H^+]$ and the $[OH^-]$ in a solution of 0.40 g of NaOH dissolved in 10.0 cm^3 of solution.

12. The pH is used to measure the acidity of a solution, where **$pH = -\log[H^+]$**. At 298K, if a solution's:

- pH < 7, the solution is acidic
- pH = 7, the solution is neutral
- pH < 7, the solution is basic

CQD 10. [G] If the pH of a solution is 5.00, what is $[H^+]$? Is the solution acidic or basic?

13. At 298K, a solution is acidic if its pH < 7, basic if its pH > 7 and neutral if its pH = 7.

CQD 12. [G] Classify the following solutions as acidic, basic or neutral.

a. a solution of pH = 5
b. a solution of pH = 8
c. a solution of pH = 7

CQD 18. Calculate the $[H^+]$ and $[OH^-]$ in a solution whose:
a. pH = 5.0
b. pH = 8.0
c. pH = 7.0
d. Are the above solutions acidic, basic or neutral?

CQD 21. [G]What is $[OH^-]$ in a solution whose pH is 8? Is the solution acidic or basic?

14. Characteristic properties of acids include their reactions with:

- Metals to produce hydrogen gas, metal + acid $\rightarrow$ salt + hydrogen gas
  Hydrohalic acids produce metal halides
  Nitric acid produces metal nitrate
  Sulfuric acid produces metal sulfates

  $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
  magnesium chloride
  $Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
  magnesium nitrate
  $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
  magnesium sulfate
  $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
- Bases to produce salts, acid + base $\rightarrow$ salt (+ water)
  $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
  Reactions between acids and bases are called neutralization reactions. When a strong acid reacts with a strong base, the net ionic neutralization reaction is always:
  $H^+(aq) + OH^- \rightarrow H_2O(l)$
- Carbonates to produce carbon dioxide gas,
  carbonate + acid $\rightarrow$ salt + carbon dioxide gas + water
  $MgCO_3(s) + 2HCl \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$
  $MgCO_3(s) + 2H^+ \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

Ex. 2 Acids react with bases, carbonates and metals. Which of the following reactions produces gases?

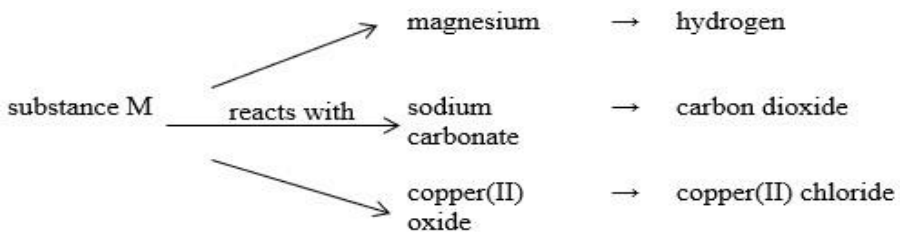
	base	carbonate	metal
[-A-]	√	√	√
[-B-]	√	×	×
[-C-]	×	√	√
[-D-]	×	√	×
[-E-]	×	×	√

CQD 13. Write the reaction between:

a. HCl(g) and water.
b. HNO_3 and water.
c. H_2SO_4 and water. Show both steps of dissociation.
d. $CH_3COOH(aq)$ and water.

CQD 31. The reaction of a strong acid with a strong base to produce salt and water is called:

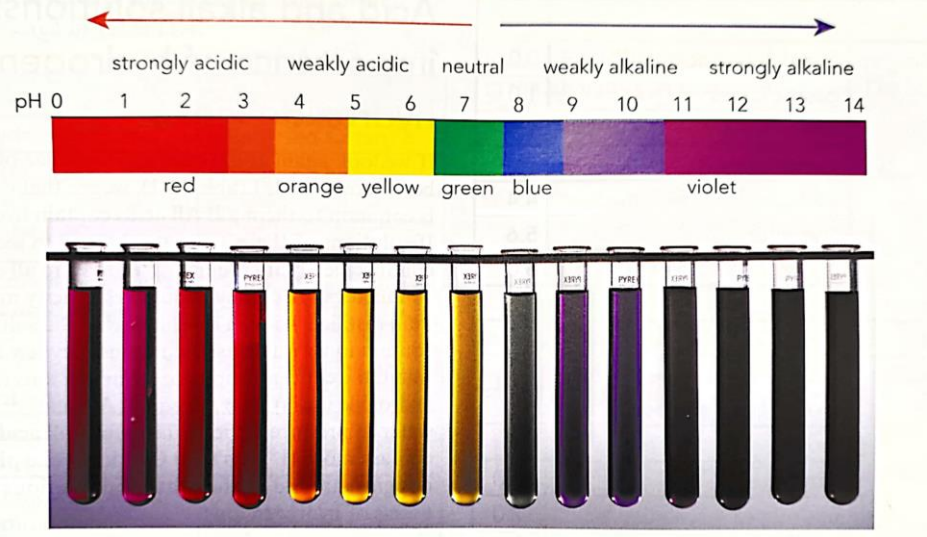
CQD37. Some reactions of substance M are shown below.



15. An indicator is a substance that changes color when placed in acid or base.
16. Effect of acids and bases on indicators

	Color in		
Indicator	Acid	Neutral	Base
Litmus	**red**	**purple**	**blue**
Thymolphthalein	colorless	colorless	**blue**
Methyl orange	**red**	**orange**	**yellow**
phenolphthalein	colorless	colorless	**pink**

17. pH can be measured using a pH-meter or a universal indicator paper. When pH is measured using the universal indicator paper, it should be reported using whole numbers only (rough).
18. A Universal indicator is a mixture of dyes that is very useful as it gives a range of colors depending on a solution's relative acidity or basicity.



From Chemistry for Cambridge IGCSE Coursework Textbook, edition 5

CQD 6. If the burette was filled with dilute hydrochloric acid, and the solution was slowly added to a conical flask containing dilute sodium hydroxide solution, what will be the color change in the conical flask if the following indicators are added?

a) Litmus indicator is used: from _____ to _____
b) Methyl orange is used: from _____ to _____
c) Thymolphthalein is used: from _____ to _____
d) Phenolphthalein is used: from _____ to _____

19. Universal indicator is a dye that changes color depending on the acidity or basicity of the solution.

Ex. 9 When a drop of universal indicator is added to acetic acid solution, the clourless liquid turns:

a) pink
b) blue
c) orange
d) green
e) purple

When a drop of universal indicator is added to sodium hydroxide solution, the clourless liquid turns:

a) red
b) yellow
c) orange
d) green
e) purple

When a drop of universal indicator is added to water, the colorless liquid turns:

a) red
b) yellow
c) orange
d) green
e) purple

20. Bases are oxides or hydroxides of metals and **alkalis** are soluble bases.

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

$NH_3(aq) + H_2O(l) \rightleftarrows NH_4^+(aq) + OH^-(aq)$

21. Characteristic properties of bases include their reactions with:

- Acids to produce salts, acid + base → salt (+ water)

  $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

  $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

  $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$

  $CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

  $CaO(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l)$

- Ammonium salts, ammonium salt + base → ammonia + salt + water

  $NH_4NO_3(s) + KOH(aq) \rightarrow NH_3(g) + KNO_3(aq) + H_2O(l)$

  $NH_4NO_3(s) + OH^-(aq) \rightarrow NH_3(g) + NO_3^-(aq) + H_2O(l)$

  $2NH_4NO_3(s) + CaO(s) \rightarrow 2NH_3(g) + Ca(NO_3)_2(s) + H_2O(l)$

CQD30. Which of the following reactions produce(s) a gas?

a) acid + base
b) ammonium salt + base
c) acid + metal carbonate
d) acid + metal
e) acid + metal sulfite
f) acid + metal oxide

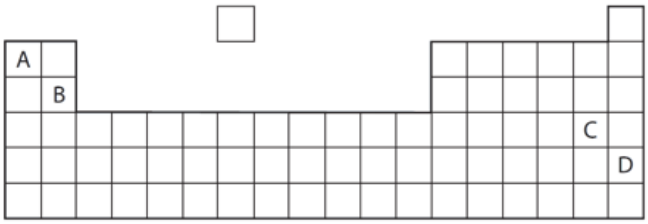
22. Metallic oxides, except Al_2O_3, ZnO, PbO, are usually basic oxides. They react with acids to give salt and water, for eg. CuO and CaO.

23. Non-metallic oxides are usually acidic oxides. They react with bases to give salt and water, for eg. SO_2 and CO_2.

24. Amphoteric oxides react with both acids or bases to produce salt and water. Examples include: ZnO, PbO and Al_2O_3.

25. Neutral oxides do not react with acids or bases. Examples include some non-metallic oxides like H_2O, N_2O, CO.

CQD 4. Which of the following elements below is most likely to form an acidic oxide?



A. A only
B. B only
C. C only
D. D only
E. A, B, and C

CQD 2. Consider the following.

Aluminum oxide	Al_2O_3
Calcium oxide	CaO
Carbon dioxide	CO_2
Carbon monoxide	CO
Magnesium oxide	MgO
Sulfur dioxide	SO_2

Which of the oxides listed above

i. can react with hydrochloric acid but not with aqueous sodium hydroxide?
ii. can react with aqueous sodium hydroxide but not with hydrochloric acid?
iii. can react with both hydrochloric acid and aqueous sodium hydroxide?
iv. cannot react with hydrochloric acid or aqueous sodium hydroxide?

CQD 3 [G] Fill in the table to identify each of the given oxides as acidic, basic, neutral, or amphoteric.

Oxide	**Acid-base nature**
MgO	
CO_2	
CuO	
N_2O	
CO	
ZnO	
SO_2	
Al_2O_3	
CaO	

Ex. 4 Element X is on the left-hand side of the Periodic Table. Which of the following is true about its oxide?

	Type of oxide	Nature of oxide
[-A-]	metallic	acidic
[-B-]	metallic	basic
[-C-]	non- metallic	acidic
[-D-]	non- metallic	basic
[-E-]	metallic	neutral

CQD 14. Complete the equations:

a. $CaO(s) + H_2O(l)$
b. $SO_3(g) + H_2O(l)$
c. $CO_2(g) + H_2O(l)$
d. $SO_2(g) + H_2O(l)$
e. $K_2O(s) + H_2O(l)$
f. $Na_2O(s) + H_2O(l)$

CQD 22. [T] An oxide **X** reacts with both potassium hydroxide and sulfuric acid. Which of the following is true about **X**?

A. **X** is an acidic oxide.
B. **X** is a basic oxide.
C. X is an amphoteric oxide.
D. **X** is a neutral oxide.
E. none of the above**.**

CQD 20. [G] A solution of concentrated hydrochloric acid, HCl, is added to an aqueous solution of sodium carbonate, Na_2CO_3.

a. What is the expected observation?
b. Give the net ionic equation for the reaction.

CDQ 15. Write down the complete acid-base equations of the following:
a. Neutralization of an acid and a base
b. Carbonate ion in water
c. Sodium carbonate with a strong acid (net ionic reaction)
d. Ammonium ion in water
e. Strong base and ammonium salt
f. $CaO(s) + 2H^+(aq)$
g. $SO_3(g) + NaOH(aq)$ (net ionic reaction)

Salt Preparation

A salt is the product of the reaction between an acid and a base.
Different techniques can be used to prepare different salts in the laboratory.

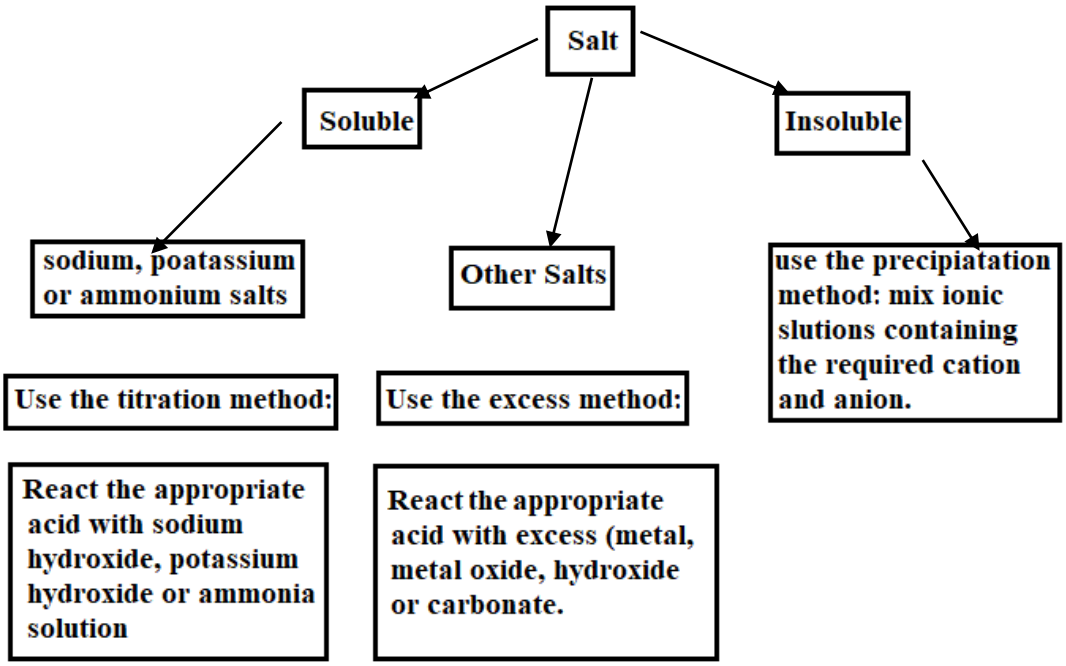
To prepare soluble salts

A. Titration method: reaction of acid and alkali
B. Excess method: reaction of acid with:
   - excess insoluble base
   - excess metal
   - excess metal oxide
   - excess metal carbonate

To prepare insoluble salts

Precipitation method

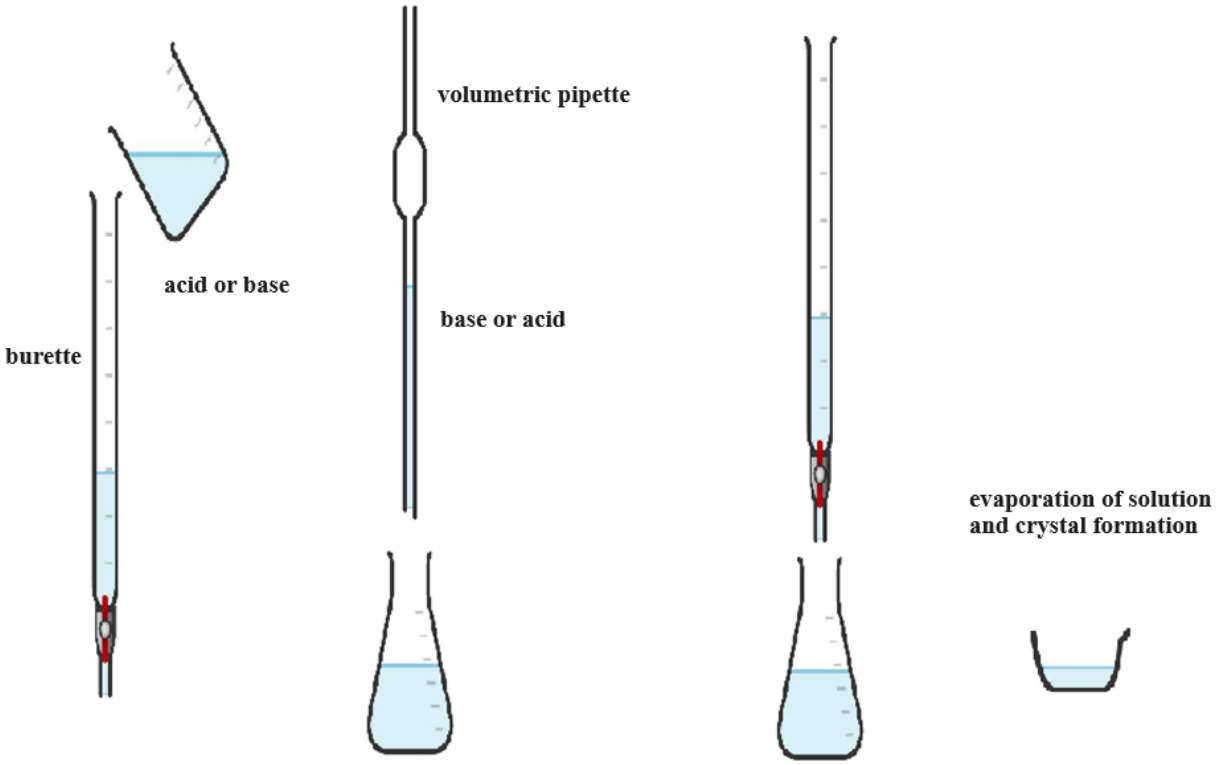
The following is a helpful summary when to use each method:


Salt
Soluble
Insoluble
sodium, poatassium or ammonium salts
Other Salts
use the precipiatation method: mix ionic slutions containing the required cation and anion.
Use the titration method:
Use the excess method:
React the appropriate acid with sodium hydroxide, potassium hydroxide or ammonia solution
React the appropriate acid with excess (metal, metal oxide, hydroxide or carbonate.

Irrespective of method, salt crystals are prepared by evaporation and crytallisation.
Remember the general solubility rules for salts:

- sodium, potassium and ammonium salts are soluble
- nitrates are soluble
- chlorides are soluble, except lead and silver
- sulfates are soluble, except barium, calcium and lead
- carbonates are insoluble, except sodium, potassium and ammonium
- hydroxides are insoluble, except sodium, potassium, ammonium and calcium (partially)

26. <u>Titration method:</u>
General set up:



27. **Titration:** is the progressive addition of a base to an acid or vice versa.
28. **Equivalence point:** The *equivalence point*, or *<u>stoichiometric</u> point*, of a chemical reaction when a titrant is added and is stoichiometrically equal to the amount of moles of substance (known as analyte) present in the sample: the smallest amount of titrant that is sufficient to fully neutralize or react with the analyte.
29. During a titration, the volume of titrant is obtained taking the difference between the last and first reading of the burette.
30. At the equivalence point of an acid-base titration, number of moles of H^+ = number of moles of OH^-

<u>Typical Question:</u>
Describe the preparation of sodium chloride from hydrochloric acid.
Include in your account details of any apparatus needed and name any other chemicals needed.

<u>Steps:</u>

- Fill a burette with aqueous sodium hydroxide.
- Read the burette.
- Using a pipette, add 25.0 ml of the aqueous hydrochloric acid into a conical flask.
- Add 2-3 drops of indicator.

- Add aqueous sodium hydroxide to the hydrochloric acid solution, while continuously swirling, till the end point is reached (point at which colour of the indicator changes upon adding a single drop of base.
- Read the burette.
- Repeat the experiment until you get concurrent results (readings that do not differ by more than 0.20 ml).
- Take the average of three concurrent readings. This will be the average volume of the base needed to neutralize the acid.
- Repeat the experiment one last time, by adding the calculated average volume of the base without using an indicator.

N.B. Another way to know when the neutralization point has been reached is to measure pH of the medium using a pH meter. The pH will be 7.00 at neutralization point.

Crystallize the salt:

- Heat the solution in an evaporating dish directly on a flame till most of the liquid evaporates.
- Transfer the evaporating dish over a steam bath. Heat gently till the crystallization point (till solution is saturated).

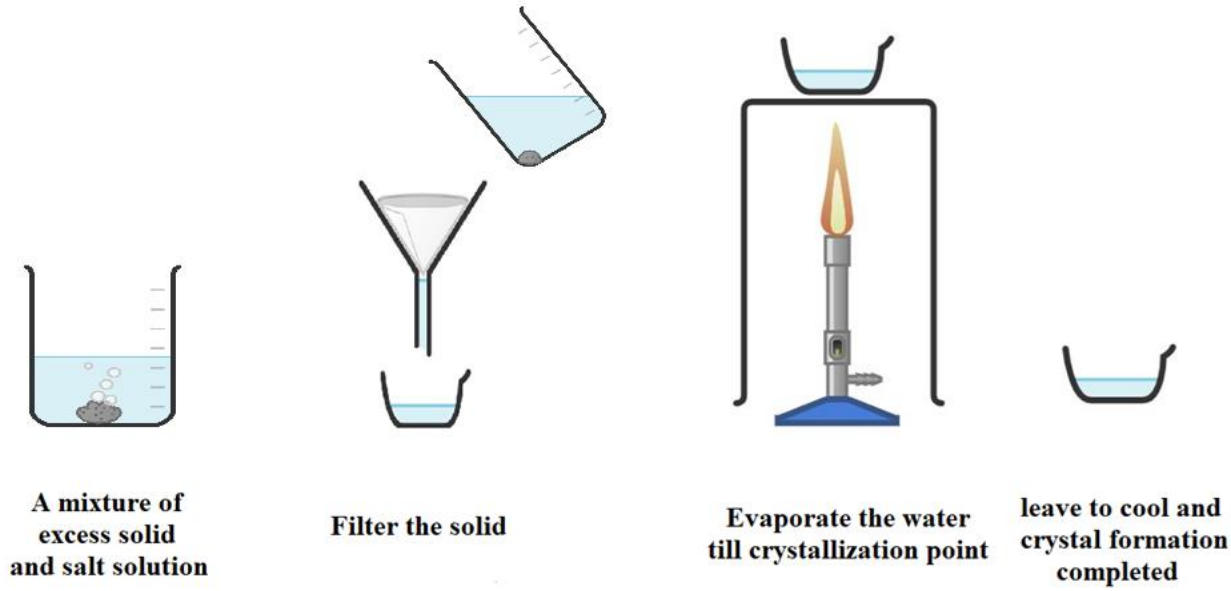
Crystallization point(saturation) is checked we can check if the solution is saturated by dipping a cold glass rod in the solution from time to time and checking for the formation of crystals.

- Leave to cool.
- Filter off the crystals.
- Dry the crystals between two filter papers without pressing or in open air.

CQD8. 200 mL of 1.00 M NaOH is added to 200 mL of 1.00 M HCl. Calculate the final $[H^+]$ and $[OH^-]$.

31. Excess method:

General set up:


A mixture of excess solid and salt solution
Filter the solid
Evaporate the water till crystallization point
leave to cool and crystal formation completed

Step 1: warm acid and turn burner off, then either

- Add excess metal till no more bubbles form, or
- Add excess metal oxide till the solution no longer turns blue litmus paper red, or
- Add excess carbonate till no more bubbles form

Step 2: filter

Step 3: crystallize

Typical Question

Describe how a solution of zinc sulphate could be made from the insoluble compounds of zinc.

Method I: Acid + Metal → Salt + Hydrogen

Used for any metal that is more reactive than hydrogen (not suitable for the preparation of Cu or Ag salts)

Steps:

- In a beaker, add a known volume of aqueous sulphuric acid.
- Gently heat the acid in the beaker on a Bunsen burner (to speed up the rate of the reaction)
- Turn off the burner. (To prevent splashing of the liquid).
- Start adding spatula measures of the zinc metal to the acid and stir.
- Continue adding the metal till:
    - Excess metal settles at the bottom of the beaker AND
    - No more bubbles of hydrogen gas evolve.
- Filter, the residue is excess unreacted metal and the filtrate is zinc sulphate solution.
- Crystallize the salt.

Method II: Acid + Metal Oxide → Salt + Water
Used for any metal compound even Cu and Ag salts.

Steps:

- In a beaker, add a known volume of aqueous sulphuric acid.
- Gently heat the acid in the beaker on a Bunsen burner (to speed up the rate of the reaction) ***throughout the experiment.***
- Start adding spatula measures of zinc oxide to the acid and stir.
- Continue adding the metal oxide till excess metal oxide settles at the bottom of the beaker.
- Filter, the residue is excess unreacted metal and the filtrate is zinc sulphate solution.
- Crystallize the salt.

Method III: Acid + Metal Carbonate → Salt + Water + Carbon dioxide
Used for any metal compound even Cu and Ag salts.

Steps:

- In a beaker, add a known volume of aqueous sulphuric acid.
- Start adding spatula measures of zinc carbonate to the acid and stir.
- Continue adding the metal carbonate till:
    - Excess metal carbonate settles at the bottom of the beaker AND
    - No more effervescence is seen (due to evolution of CO_2 gas).
- Filter, the residue is excess unreacted metal and the filtrate is zinc sulphate solution.
- Crystallize the salt.

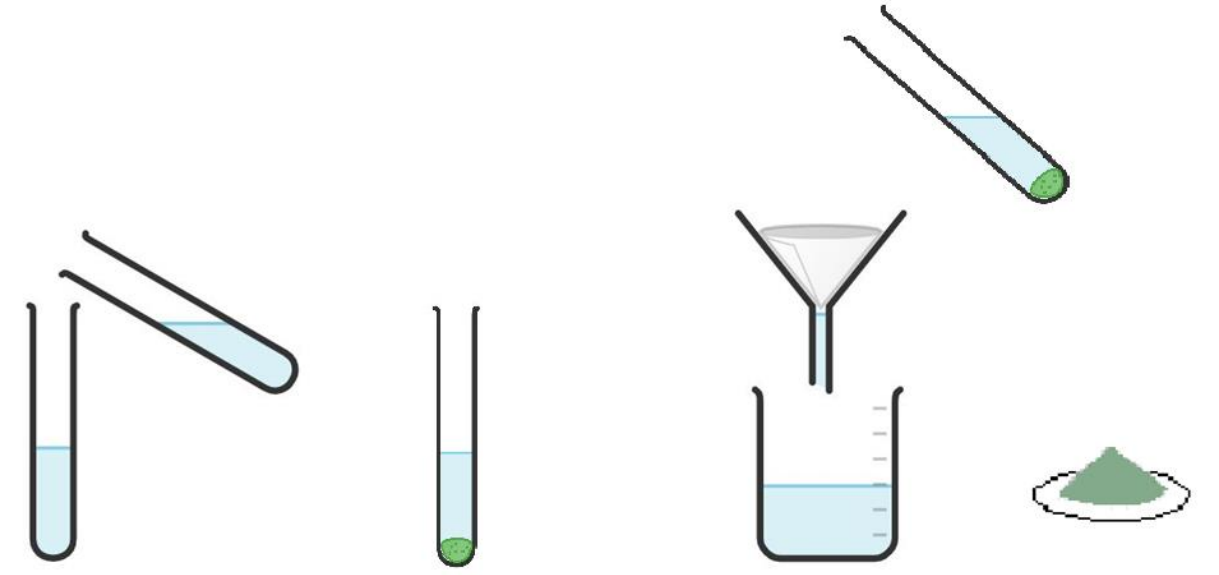
The volume of acid used in excess method need not be specified. The solid (metal, metal oxide or metal carbonate) is added till in excess to make sure that all the acid has been used up (or reacted or consumed) irrespective of the volume of the acid used.

32. Summary:

Reagent	Procedure	How to know reaction is complete (All acid has reacted)
Acid + Metal	- Heat acid first. - Put off Bunsen burner. - Start adding spatula measures of metal.	- Excess metal remains at the bottom of the beaker - No more effervescence
Acid + Metal oxide	- Continue heating whilst adding the metal oxide to the acid.	- Excess metal oxide remains at the bottom of the beaker.
Acid + Metal carbonate	- No heating at all as effervescence occurring is vigorous (to prevent splashing of the liquid)	- Excess metal carbonate remains at the bottom of the beaker - No more effervescence.

33. Precipitation method:

General set up:



Solution A (containing reuired cation) is mixed with solution B (containing required anion) are mixed

precipitate is formed

filter, wash residue in little cold water

damp filter paper with salt left to dry or placed in oven

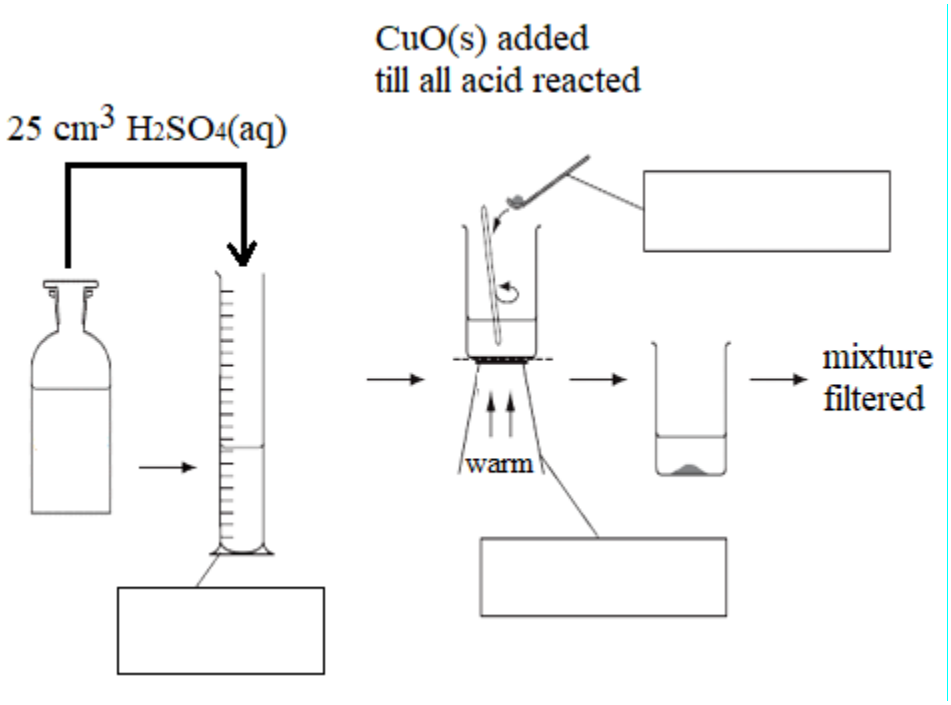
Typical question

How can a pure sample of the insoluble salt, lead (II) sulphate, be made?

Ex. 5 Which of the following is the best method to prepare $ZnSO_4$ from H_2SO_4 and ZnO?

a) titrate ZnO(s) against H_2SO_4.
b) Mix equal amounts of the two reagents and crystallize.
c) Add ZnO to excess acid solution.
d) Add excess ZnO to acid solution.
e) All of the above work.

Ex. 6 Copper sulphate solution was prepared by reacting *excess* copper oxide with dilute sulphuric acid. The diagram below shows the method used.



a) Fill in the empty boxes.
b) Explain what is meant by excess.
c) Draw a diagram to show how the mixture is filtered.

Ex. 7 The pieces of apparatus needed during the preparation of sodium chloride, NaCl, by **titration** are:

A. stop clock, burette, dropper
B. balance, burette, beaker
C. flask, burette, pipette
D. pipette, balance, burette
E. pipette, stop clock, balance

CQD 5.

Three methods are used for salt preparation

Method A: using a burette and an indicator.
Method B: mixing two solutions to obtain a salt by precipitation.
Method C: addition of an excess base or metal to a dilute acid then removing the excess by filtration.
Method D: addition of an excess metal oxide to a dilute acid then removing the excess by filtration.
Method E: addition of an excess metal carbonate to a dilute acid then removing the excess by filtration.

For the following salt preparations, choose the method name any reagent needed and complete or write the reactions.

a) The soluble salt, zinc sulfate from the insoluble base zinc oxide.

Method**:**
Reagent**:**
Word equation**:**

b) The soluble salt, potassium chloride, from the soluble base, potassium hydroxide.

Method**:**
Reagent**:**
Equation:

c) The insoluble salt, Lead(II) iodide from the soluble salt Lead(II) nitrate.

Method**:**
Reagent**:**
Equation**:**

d) Name the chemical needed to be added to insoluble magnesium oxide to form magnesium sulphate.
e) Which method can be used to prepare the soluble salt potassium chloride from potassium hydroxide?
f) Name the chemical needed to be added to lithium hydroxide to form lithium sulphate.
g) Which method can be used to prepare the insoluble salt silver chloride from the soluble salt silver nitrate?
h) Name the chemical added to silver nitrate to form silver chloride.
i) Which method can be used to prepare the soluble salt magnesium sulphate from the insoluble base magnesium oxide?
j) Name the chemical needed to be added to insoluble calcium oxide to form calcium nitrate.
k) Which method can be used to prepare the soluble salt sodium iodide from sodium hydroxide?
l) Name the chemical needed to be added to sodium hydroxide to form sodium chloride.

m) Which method can be used to prepare the insoluble salt lead (II) iodide from the soluble salt lead (II) nitrate?
n) Name the chemical added to lead (II) nitrate to form lead (II) iodide.

CQD 26. Which of the following reagents can be added to hydrochloric acid, HCl, to prepare zinc chloride, $ZnCl_2$?

a) Zn
b) ZnO
c) $ZnCO_3$
d) $ZnSO_4$

CQD 27. Which of the following reagents can be added to sulphuric acid, H_2SO_4, to prepare copper (II) sulphate, $CuSO_4$?

a) Cu
b) CuO
c) $CuCO_3$
d) $Cu(OH)_2$

CQD 28. Which of the following reactions produce ammonia gas?

a) acid + base
b) ammonium salt + base
c) acid + metal carbonate
d) acid + metal
e) acid + metal sulfide

CQD 29. Which of the following reactions produces a pungent smelling gas that turns damp red litmus blue?

a) $CaO + H_2SO_4$
b) $Ca(OH)_2 + H_2SO_4$
c) $Ca + H_2SO_4$
d) $Ca(OH)_2 + (NH_4)_2SO_4$
e) $CaCO_3 + H_2SO_4$

CQD 30. Which of the following reactions produce (s) a gas?

a) acid + base
b) ammonium salt + base
c) acid + metal carbonate
d) acid + metal
e) acid + metal sulfite
f) acid + metal oxide

CQD 32. During preparation of silver iodide, AgI, by precipitation, what is the correct order of the steps?

a) mixing two aqueous solutions
b) filtration
c) stirring with a glass rod
d) drying residue in an oven

CQD 33. The pieces of apparatus needed during the preparation of sodium chloride, NaCl, by **titration** are:

A. stop clock, burette, dropper
B. balance, burette, beaker
C. flask, burette, pipette
D. pipette, balance, burette
E. pipette, stop clock, balance

CQD 34. During the titration of an acid with a base, the color change that occurs when phenolphthalein is used as an indicator is **colorless to pink**

CQD 35. During the titration of an acid with a base, the color change that occurs when methyl orange is used as an indicator is **red to yellow**

CQD 36. Dilute sulphuric acid is added to a mixture that contains magnesium, zinc and copper in a beaker. The beaker is left for about 15 minutes and its contents are then filtered.
What does the filtrate contain? **Magnesium sulfate and zinc sulfate**

CQD 38. The information in the box is about preparation of magnesium nitrate crystals.

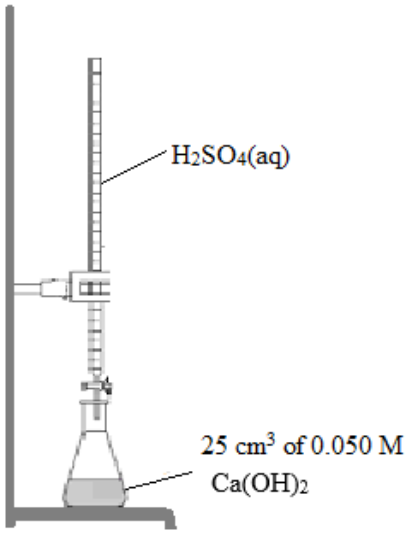
Step 1: Add spatula measures of magnesium oxide to some hot dilute nitric acid then stir.
Step 2: Continue adding magnesium oxide until it is in excess.
Step 3: Remove excess magnesium oxide so that colorless magnesium nitrate is left behind.
Step 4: Evaporate the magnesium nitrate solution until it is saturated.
Step 5: Leave the saturated solution to cool, white crystals are seen formed on cooling.
Step 6: Remove the crystals from the solution.
Step 7: Dry the crystals on filter paper.

a) Explain why excess magnesium oxide is used in step 2.
b) Explain how excess magnesium oxide is removed from the solution in Step 3.
c) What is meant by the term saturated solution?
d) What is the practical method that could show that the solution is saturated?
e) Explain why the crystals in Step 7 are dried using filter paper and not by heating.

34. Know that Cl_2 and NO_2 dissolve in water forming two acids

CQD 23. [T]

a. A student titrates an aqueous sulfuric acid solution using 25 cm^3 of 0.050 M solution calcium hydroxide and few drops of an acid-base indicator.
The volume of sulfuric acid needed to react completely with the calcium hydroxide solution is 12.5 cm^3.



The equation of the reaction is shown:

$$H_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow CaSO_4(s) + 2H_2O(l)$$

What is the concentration of sulfuric acid solution?

- Calculate the number of moles of calcium hydroxide.
- Calculate the number of moles of sulfuric acid to react completely with calcium hydroxide.
- Calculate the concentration of the sulfuric acid solution.

b. After titration is done, the final solution in the beaker contains solid calcium sulfate, water, and some dissolved acid-base indicator.
Describe how to extract dry sample of solid calcium sulfate from the solution. Write the process including all key steps.

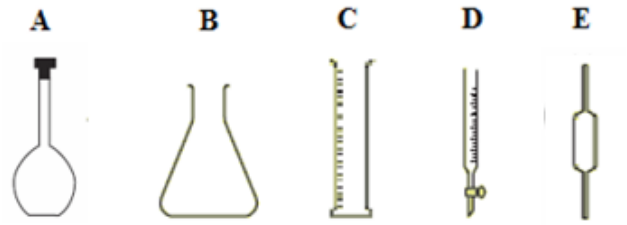
c. The student conducted the following tests on the **products** of the reaction between sulfuric acid and sodium carbonate.
Describe the student's observation in each test.
i. Flame test on the solution.
ii. Passing gas released in lime water.
iii. Addition of aqueous solution of barium nitrate.

d. Dilute sulfuric acid reacts with metals, metal oxides and bases.

Write chemical equations for the reaction of dilute sulfuric acid with each of the following:
i. magnesium.
ii. sodium oxide
iii. potassium hydroxide

CQD 24. [T]

A student conducts an acid-base titration to determine the concentration of a sodium hydroxide solution using a 0.1 M hydrochloric solution. He is provided with the following apparatus.



a. Name the items that the student needs to perform the titration.

b. The student adds few drops of an acid-base indicator to 20 cm^3 of sodium hydroxide and titrates the mixture with the hydrochloric acid solution.
i. Why is it necessary to add an acid-base indicator before starting the titration?

ii. The student makes two reading to determine the volume of the hydrochloric acid needed to neutralize completely the sodium hydroxide solution. The initial reading he made before starting the titration is 10 cm^3 and the final reading at the end of the titration is 23 cm^3.
What is the volume of hydrochloric acid used to completely neutralize 20 cm^3 of sodium hydroxide solution?

c. The student repeats the same experiment using the same materials. He discards the content of the final solution from the flask and refills it directly with another 20 cm^3 of the same sodium hydroxide.
What did the student do wrong in this procedure? Describe the correct procedure.

CQD 24. [T] Plan an experiment to find the concentration of an aqueous solution of sulfuric acid provided with dilute solution of sodium hydroxide of known concentration and common laboratory apparatus.

Chapter 4-Oxidation-Reduction Reactions

1. A combustion reaction is defined as a reaction with a gas resulting in the production of a large amount of heat and light. Typical such reactions are reactions with oxygen. For example, benzene burns in oxygen as shown by the reaction below.

$$2C_6H_6(g) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

2. Not all reactions with oxygen release a large amount of heat and light, for example the reaction with copper metal.

$$Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

3. An oxidation can be defined as the gain of oxygen and reduction as the loss of oxygen. During a combustion, the substance reacting with oxygen is being oxidized. In the above reaction, copper metal is being oxidized. In both reactions, oxygen is said to be reduced. Oxidations and reduction reaction occur together, never one alone, this is why it is best to refer to such reactions by the term redox reactions.

4. Redox reactions need not involve oxygen gas. Consider the two reactions below:

$CuO + H_2 \rightarrow Cu + H_2O$ CuO lost O-atom, so it is reduced, H_2 gained O-atoms so it is oxidized

$ZnO + C \rightarrow Zn + CO$ ZnO lost O-atom, so it is reduced, C gained O-atoms so it is oxidized

5. Everyday life is full of redox reactions, two example:

- Corrosion: metals are attacked by oxygen and water changing them into the metal oxide, most famous is rusting of iron
- When food (organic matter) goes bad, it is attacked by oxygen and we say food turned rancid.

6. **Demonstration:** place a zinc rod in an aqueous copper(II) solution**.**

7. **Demonstration:** place a copper strip in an aqueous zinc solution.

8. The reaction occurring is called a *displacement* reaction (a reaction in which a more reactive element displaces a less reactive one from a solution of its salt).

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

9. In the above reaction, Zn lost 2 electrons and Cu^{2+} gained them. We can write the reaction as two *half* reactions:

$$Zn \rightarrow Zn^{2+} + 2e^-$$
$$Cu^{2+} + 2e^- \rightarrow Cu$$

10. As such we can define redox reactions in terms of electrons:
Oxidation: is a half reaction in which electrons are lost.
Reduction: is a half reaction in which electrons are gained.

11. **Reducing agent:** species that reduces another species and gets itself oxidized**.**

12. **Oxidizing agent:** species that oxidizes another species and gets itself reduced.

Ex. 1 In the reaction: $Fe^{+3} + Cu^{+} \rightarrow Fe^{+2} + Cu^{+2}$
Fe^{+3} is a reducing agent and Cu^{+} is an oxidizing agent.

a) Fe^{+3} is an oxidizing agent and Cu^{+} is a reducing agent.
b) Fe^{+2} is an oxidizing agent and Cu^{+2} is a reducing agent.
c) Fe^{+3} is oxidized and Cu^{+} is reduced.
d) An oxidizing agent is itself a reducing agent.

Ex. 2 A spontaneous reaction occurs when a piece of iron is placed in 1.0M $CuSO_4$ (aq). The reducing agent is:

a) Fe.
b) Cu^{+2}.
c) H_2O.
d) SO_4^{-2}.
e) $CuSO_4$.

13. To keep track of electrons in redox reactions can be difficult. Look at the reaction below:

$$8H_2O + Cr_2O_7^{2-} + 3CH_3OH \rightarrow 3CH_2O + 2Cr^{3+} + 7H_2O + 8OH^-$$

Which species is gaining and which is losing electrons?
To answer this question, scientists came up with a fictious charge they called oxidation state which allows keeping track of electrons in redox reactions.
The rules to assign oxidation numbers are:

The oxidation state of

- an element is 0
- a simple ion is its charge
- fluorine in its compounds is -1
- hydrogen in its compounds is +1 except in metallic hydrides in which it is -1
- oxygen in most of its compounds is -2
- aluminium in its compounds is +3
- Group I metals in their compounds is +1
- Group II metals in their compounds is +2
- chlorine, bromine and iodine in their compounds is generally -1
- Sum of oxidation numbers must be equal to the overall charge
- The sum of the oxidation numbers is zero for a compound and equal to the charge on an ion

CQD 10. [G] Find the oxidation number for each element in following species.

a. NH_3
b. H_2CO_3
c. I_2
d. MnO_4^-
e. $Cr_2O_7^{2-}$

Going back to our equation and assigning oxidation numbers we get:

$$8H_2O + Cr_2O_7^{2-} + 3CH_3OH \rightarrow 3CH_2O + 2Cr^{3+} + 7H_2O + 8OH^-$$
$$\textbf{+1 -2 \quad +6 -2 \quad -2 +1 -2 +1 \quad 0 +1 -2 \quad +3 \quad +1 -2 \quad -2 +1}$$

We note that the oxidation number of Cr decreased, so $Cr_2O_7^{2-}$ is reduced $\Rightarrow$ it is the oxidizing agent
We note that the oxidation number of C increased, so CH_3OH is oxidized $\Rightarrow$ it is the reduing agent

14. We can add to the definitions of oxidation and reduction:
- **Oxidation:** an increase in oxidation number.
- **Reduction:** a decrease in oxidation number

15. So a reaction can be identified as a redox reaction if:
- It involves gain and loss of O-atoms
- Loss and gain of electrons

16. For some reactions, a change in color can also identify redox reactions. For eg.,
- Whenever a reaction involves acidified aqueous $KMnO_4$ (potassium manganate(VII)), which is purple and turn colorless.
- Whenever a reaction involves aqueous potassium iodide which is colorless and turns brown.

CQD 19.
a. When acidified potassium manganate (VII) is added to a reducing agent its color changes from purple to _____.

b. When acidified potassium iodide solution, which is a reducing agent, is added to an oxidizing agent such as chlorine or hydrogen peroxide, the solution turns a _____ color.

CQD1. [G] Which of the following equations is an oxidation-reduction (redox) reaction?

a. $CuCO_3 \rightarrow CuO + CO_2$
b. ZnO + 2HCl → ZnCl2 + H2O
c. $2Cu + O_2 \rightarrow 2CuO$
d. $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

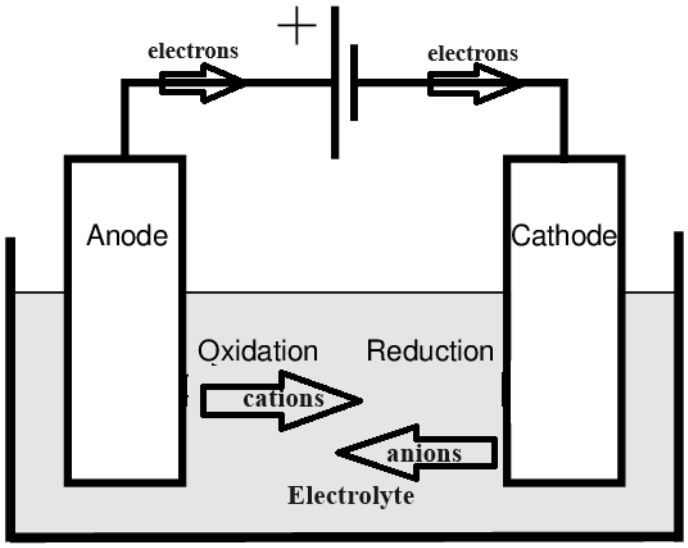
Ex. 3. Identify which of the following reactions is a redox reaction.
For the redox reactions, identify the oxidizing and reducing agents.

a. $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
b. $2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$
c. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
d. $2Fe_2O_3(s) + 3C(s) \rightarrow 3CO_2(g) + 4Fe(s)$
e. $2HCl(aq) + Zn(s) \rightarrow H_2(g) + ZnCl_2(aq)$
f. $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

17. **Electrolysis:** is the decomposition of an ionic compound, when molten or in aqueous solution, by the passage of an electric current.

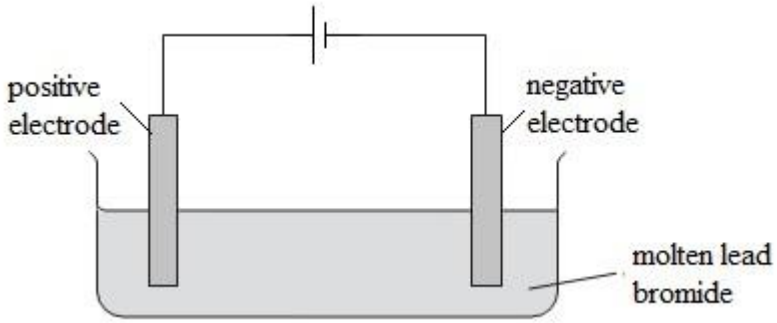
18. Below is the diagram of a simple electrolytic cell.

- the anode is the positive electrode that attracts anions and at which oxidation takes place**.**
- the cathode is the negative electrode which attracts cations and at which reduction takes place.
- The electrolyte is the molten or aqueous substance that undergoes electrolysis.
- Electrons flow from the anode to the cathode



19. During the electrolysis of a molten salt: the metal is produced at the cathode (negative electrode) and the non-metallic element at the anode (positive electrode). The electrodes used are inert (usually platinum or graphite/carbon)

CQD 11. [G] Electrolysis of molten lead bromide, $PbBr_2$, is conducted using the below apparatus.



a. Write the half-reaction that takes place at the negative electrode (cathode).
b. Write the half-reaction that takes place at the positive electrode (anode).
c. Write the chemical equation for the electrolysis of molten lead bromide.

Ex 4. What will be observed at each electrode when molten lead(II) bromide is electrolyzed?

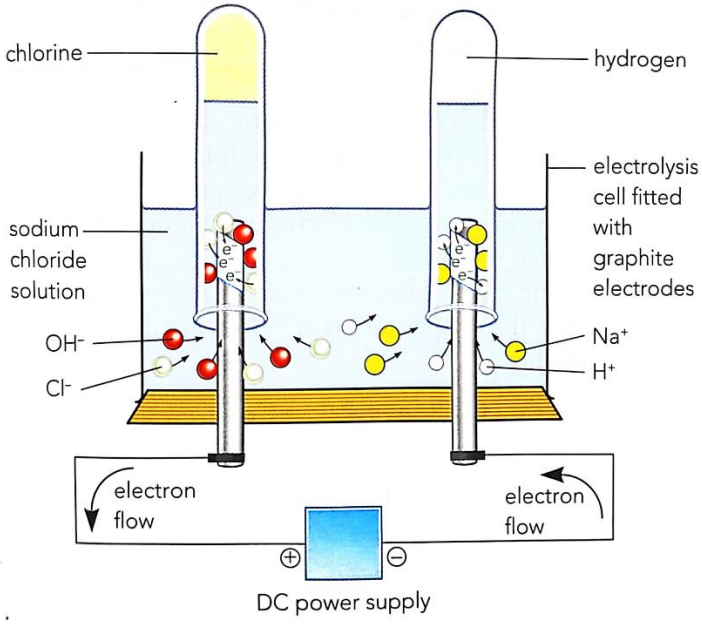
Ex. 5 Write the two half-reactions occurring when molten $ZnCl_2$ is electrolyzed.

20. Many metals such as sodium, calcium and aluminium are extracted from their ores by electrolysis of their molten salts.

21. To predict the product of electrolyzing solutions is a little more complicated. Let us discuss several examples.

Electrolysis of concentrated aqueous sodium chloride using inert electrodes

There are four ions in solution: Na^+, Cl^-, H^+ and OH^-.



From Chemistry from Cambridge IGCSE Coursebook, edition 5

- At the anode: a yellow gas is collected as chloride ions get discharged forming chlorine gas.

$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

- At the cathode: a colorless gas is collected as hydrogen ions get discharged forming hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

In the solution, we are left with Na^+ and OH^-, so the overall equation will be:

$$2NaCl(aq) + 2H_2(l) \rightarrow 2naOH(aq) + H_2(g) + Cl_2(g)$$

As similar reaction will occur if any concentrated metal halide is used: hydrogen will be produced at the cathode and the halogen at the anode. The situation will differ however if dilute electrolyte is used. Had dilute metal halide been used, then the hydroxide ion gets discharged at the anode forming oxygen gas rather than the halogen. So, if dilute NaCl(aq) is electrolyzed, hydrogen is produced at the cathode and oxygen at the anode:

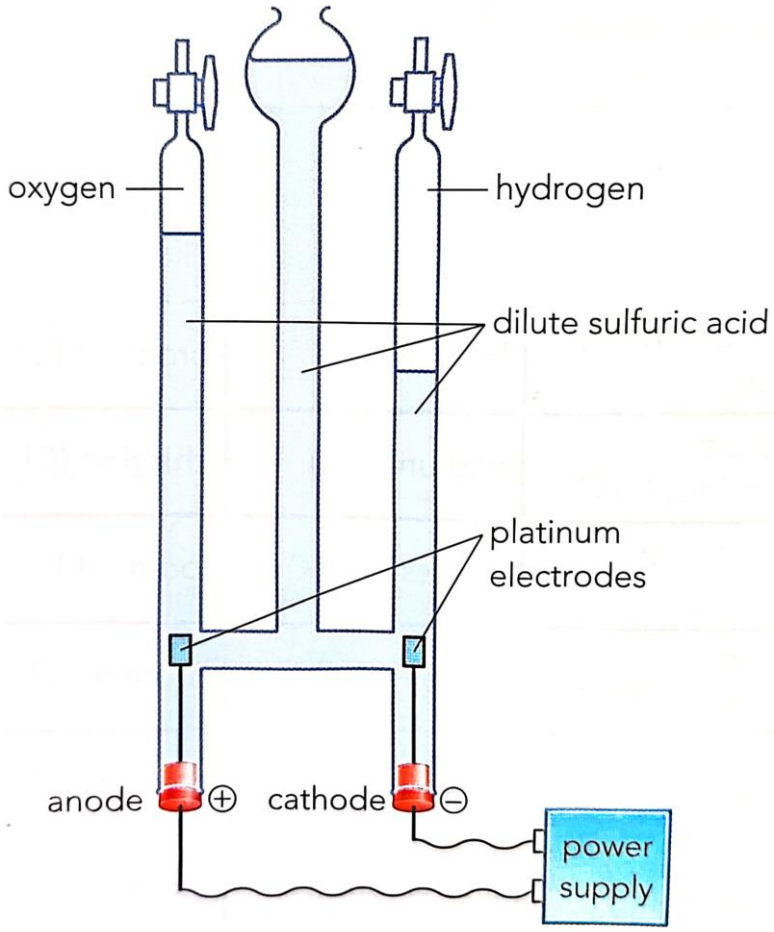
Anode: $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

Cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overall: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

Electrolysis of dilute aqueous sulfuric acid using inert electrodes

There are three ions in solution: SO_4^{2-}, H^+ and OH^-.



From Chemistry from Cambridge IGCSE Coursebook, edition 5

- At the anode: a colorless gas is collected as hydroxide ions get discharged forming oxygen gas.

$$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$$

- At the cathode: a colorless gas is collected as hydrogen ions get discharged forming hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

Note that the hydrogen and oxygen gases are produced in a ratio of 2 to 1 since the overall reaction is:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

So the concentration of the electrolyte affects the products of electrolysis, what about the nature od the electrodes?

Electrolysis of dilute aqueous copper(II) sulfate using inert electrodes

- At the cathode: a red brown metal deposits as copper(II) ions get discharged forming copper metal. The cathode increases in mass. This occurs because Cu is less reactive than H_2.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

- At the anode: a colorless gas is liberated as hydroxide ions get discharged forming oxygen gas.

$$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$$

The blue color of the solution fades and the solution becomes acidic as Cu^{2+} and OH^- are used up.

Electrolysis of dilute aqueous copper(II) sulfate using <u>copper</u> electrodes

- At the cathode: a red brown metal deposits as copper(II) ions get discharged forming copper metal. The cathode increases in mass. This occurs because Cu is less reactive than H_2.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

- At the anode: The cathode decreases in mass, as the metal gets oxidized.

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$$

The net result is a transfer of mass from anode to cathode with the solution color unchanged as the $[Cu^{2+}]$ in solution is unchanged.

General Rules:

At the cathode:

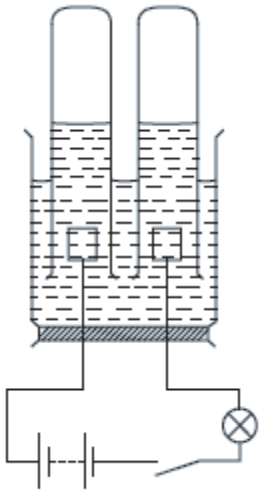
- Either a metal or hydrogen will form.
- The more reactive a metal is, the LESS likely it is to form and hydrogen forms instead. The order being from least to most likely to be discharged: Na^+ Mg^{2+} Al^{3+} Zn^{2+} H^+ Cu^{2+} Ag^+
  In other words, if H^+ and Zn^{2+} are present in a solution, H^+ gets discharged, on the other hand, if H^+ and Ag^{2+} are present in a solution, Ag^+ gets discharged.

At the anode:

- If halide ions are concentrated, they will be oxidized forming the halogen.
- If halide ions are absent or dilute, hydroxide ions will be oxidized forming oxygen.

The order being from least to most likely to be discharged: SO_4^{2-} NO_3^{2-} OH^- Cl^- Br^- I^-

CQD 13. Concentrated hydrochloric acid is electrolyzed using the apparatus below.



a. Label on the diagram the position of the electrodes.
b. What are the two observations when the circuit is switched on?
c. Give the name of the product at the positive electrode.
d. Give a test used to identify the product formed at the positive electrode.

CQD 14. Which of the following aqueous solutions produces $H_2(g)$ and $O_2(g)$ during electrolysis?

a. KI(aq)
b. $CuI_2(aq)$
c. molten NaOH
d. $CuSO_4(aq)$
e. $K_2SO_4(aq)$

	Product at anode	***Product at cathode***
a) KI(aq)	*I_2 (s)*	*H_2 (g)*
b) $CuI_2(aq)$	*I_2 (s)*	*Cu (s)*
c) molten NaCl	*Cl_2 (g)*	*Na (l)*
d) $CuSO_4(aq)$	*O_2 (g)*	*Cu (s)*
e) $K_2SO_4(aq)$	*O_2 (g)*	*H_2 (g)*

CQD 18. Fill the table with the correct products formed when aqueous copper(II) sulfate solution is electrolyzed using graphite (C) and/or copper (Cu) electrodes.

	Electrolysis of $CuSO_4(aq)$	
	Products at anode	Products at cathode
Cu electrodes		
Graphite (C) electrodes		

22. Using inert **electrodes** like carbon or platinum

The following are results for the experiments on electrolysis using inert electrodes.
Complete the table below, the first one has been done for you.

Electrolyte	**Change at inert negative electrode**	**Change at inert positive electrode**
Molten lead(II) bromide		
Concentrated aqueous sodium chloride		
Dilute sulfuric acid		
Aqueous copper (II) sulfate		

Ex. 6 The following are results for the experiments on electrolysis using inert electrodes.
Complete the table below, the first one has been done for you.

Electrolyte	**Change at the negative electrode**	**Change at the positive electrode**	**Change to electrolyte**
	Potassium formed	Iodine formed	Used up
Dilute aqueous sodium chloride			
	Hydrogen formed	Bromine formed	Potassium hydroxide formed

CQD 12. [G] Identify in each of the following the products that form on the electrode of the electrolytic cell.

Electrolyte	**At the cathode (-)**	**At the anode (+)**
Concentrated solution of sodium bromide, NaBr		
Concentrated solution of silver nitrate, $AgNO_3$		
dilute solution of potassium bromide, KBr		

23. **Electroplating:** using electricity to cover an object with a thin layer of metal**.**

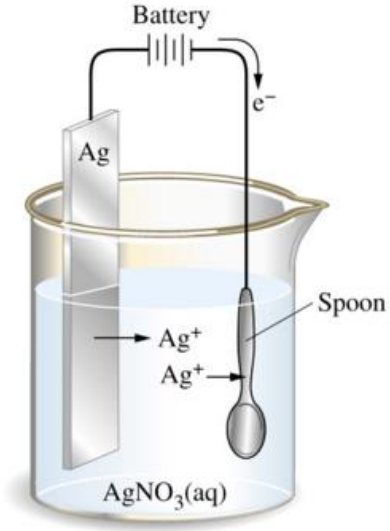
24. Metal objects are electroplated to improve their appearance and resistance to corrosion.

25. Know the apparatus needed to electroplate an object with Cu or Ag or Ni

- The object to be plated is made the cathode by connecting it to the negative terminal of the power source.
- The metal to plate is made the anode by connecting it to the positive terminal of the power source.
- Both are dipped into an electrolyte, preferable a solution of the metal used to plate.
- The metal gets oxidized and the metal ions travel to the object to be plated where they gets reduced and sticks to the object.

Electroplating a spoon with silver

- At anode: $Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$
- At cathode: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

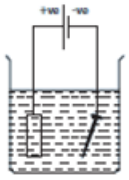


https://chem.libretexts.org/Bookshelves/Analytical_Chemistry/Supplemental_Modules_%28Analytical_Chemistry%29/Electrochemistry/Electrolytic_Cells/Electroplating

CQD 15. An iron spoon is to be plated with copper metal. Write the two half-reactions occurring at each electrode.

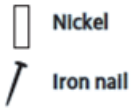
CQD 16. Which of the following apparatus could be used to electroplate an iron nail with nickel?

A

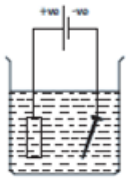


Aqueous nickel (II) sulfate

Key

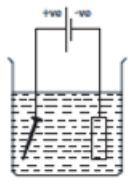


B



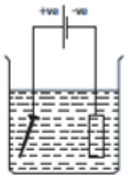
Aqueous iron (II) sulfate

C



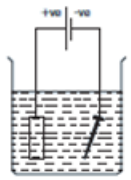
Aqueous nickel (II) sulfate

D



Aqueous iron (II) sulfate carbonate

E



nickel (II)

26. During electrolysis, electricity is used to "force" a chemical reaction that normally does not take place to take place. There are however, several reactions that occur and allow us to harvest chemical energy and change it to electrical, as in batteries chemically known as electrochemical cells. Batteries always runout and they need to be recharged or discarded. A more efficient way to change chemical energy into electrical is a fuel cell. Such cells work continuously as long as you feed them reactants. A typical fuel cell is the hydrogen-oxygen fuel cell. In this cell, hydrogen and oxygen gases react forming water as the only product and releasing a huge amount of energy. Despite its advantages as a power source, the hydrogen-oxygen fuel cells come with its own set of issues like safe storage and safe transport. Below is a table of its advantages and disadvantages.

Advantages	Disadvantages
• Renewable energy source provided it is manufactured using solar or wind energy • No emission of CO_2 • Non-toxic • Lower flammability than petrol	• Becomes non-renewable if manufactured using nuclear or fossil fuel derived energy • Requires special tanks for storage • Requires in creasing number of stations for cars to fill up • Cars need to have their engine's redesigned • Much more expensive compared to petrol

CQD 17. [G] Hydrogen fuel cells may be used to replace combustion engines in cars.

a. Give the reactants involved in the reaction occurring inside a hydrogen fuel cell.
b. Give the chemical equation for the reaction that occurs in a hydrogen fuel cell.
c. What are the advantages of hydrogen fuel cells (or hydrogen-oxygen fuel cells)?

CQD 21. [T] Which of the following electrolytic cells forms gases at both of its inert electrodes?

A. Cell 1: aqueous sodium iodide
B. Cell 2: aqueous copper(II) sulfate
C. Cell 3: aqueous sodium chloride
D. Cell 4: molten sodium chloride
E. Cell 5: molten potassium iodide

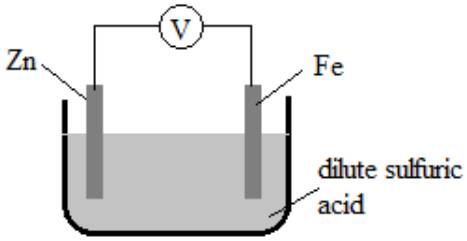
CQD 22. [T] Which of the following represent(s) a reduction change?

1. $Fe \rightarrow Fe^{2+}$
2. $H_2O_2 \rightarrow H_2O$
3. $Cr_2O_4^{2-} \rightarrow Cr_2O_7^{2-}$

A. 1 and 2 only
B. 1 and 3 only
C. 1 only
D. 2 only
E. 2 and 3 only

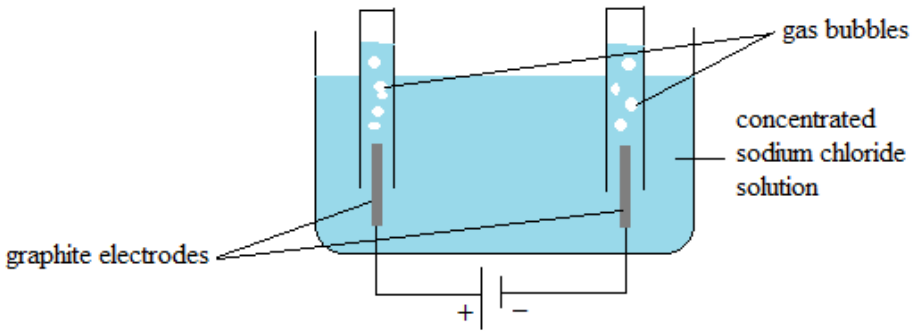
CQD 23. [T]

a. A student assembled the following electrochemical cell.



i. What is the direction of electron flow in this electrochemical cell?
ii. The student replaced the iron electrode with a silver electrode. The voltmeter reading changed. Describe the change and explain why this change occurred.
iii. When the student replaced the zinc electrode with a copper electrode, the voltmeter gave a negative reading. Explain why.

b. A student conducts electrolysis of concentrated aqueous solution of sodium chloride using the following apparatus.

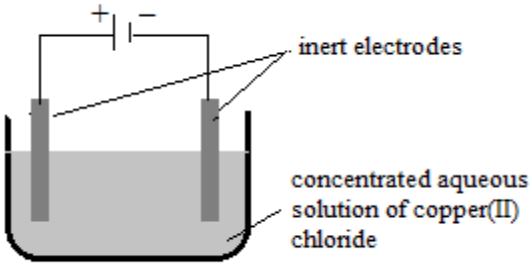


i. Which particles allows the flow of the electric current in the sodium chloride solution?
ii. Give **two** reasons for the use of graphite as the electrodes.
iii. What is the gas produced at the anode?
iv. Write the ionic half-equation for the reaction occurring at the anode. Include state symbols.
v. A blue litmus paper is dipped in the remaining solution after electrolysis is over. Does the blue litmus paper change color? Explain why.

CQD 24. [T]

Electrolysis is used to break down ionic compounds.

a. A student conducts an electrolysis for a concentrated solution of copper(II) chloride.



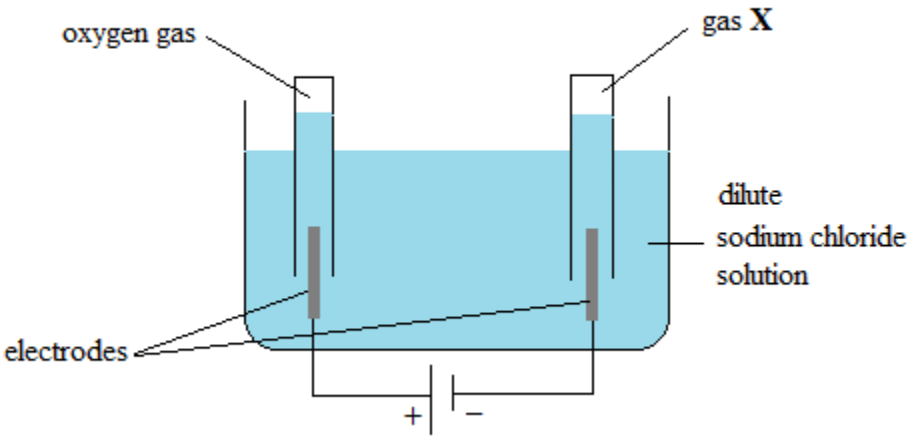
i. Why is it recommended to use inert electrodes in electrolysis?
ii. A solid deposit is observed at the negative electrode. Identify this solid.
iii. Write the ionic half equation for the reaction occurring at the negative electrode.
iv. A fizzing is observed at the positive electrode. Identify the gas released and explain how it forms.
v. Identify the species that is reduced in the solution undergoing electrolysis.

b. A student wants to electroplate an iron rod with chromium.

i. At which electrode the student should place the chromium rod?
ii. Suggest a suitable electrolyte for this electrolysis?
iii. Write the ionic half equation occurring at the negative electrode.
iv. At which electrode oxidation occurs?
v. Give a reason why chromium is used in electroplating.

CQD 25. [T]

A student conducts electrolysis on dilute solution of sodium chloride using the apparatus shown below.



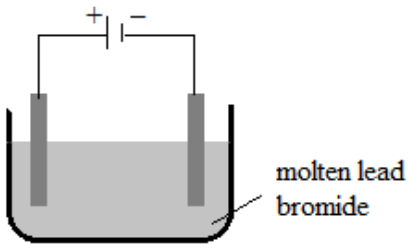
a. Gas **X** makes a pop sound when tested with a lighted splint.
Identify gas **X**.

b. Suggest a test for oxygen gas. Include result of the test in your answer.

c. Suggest a material to be used as an electrode and give one reason why it is suitable.

CQD 26. [T]

The below apparatus is used to break down molten lead bromide.



a. What is the process shown in the figure called?
b. What are the expected observations on the positive electrode (anode) and on the negative electrode (cathode)?
c. A Bunsen burner is used to melt solid lead bromide. How can the student adjust the Bunsen burner to give the hottest flame?
d. Why is it recommended to use graphite electrodes in the above apparatus?
e. When replacing molten lead bromide with an aqueous solution of lead bromide in the above apparatus, hydrogen gas is produced at the negative electrode and bromine gas is produced at the positive electrode.
   i. Suggest a test for hydrogen gas. Include observation.
   ii. How can you identify bromine gas?
f. Lead bromide is classified as hazardous chemical. Suggest a safety precaution when dealing with lead bromide.

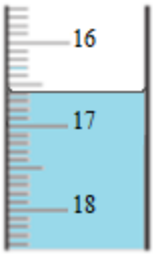
CQD 27. [T]

A student conducted two redox titration experiments to investigate the reaction between aqueous solution of potassium manganate(VII) and two aqueous solutions of iron(II) sulfate, solution **R** and solution **T**, of different concentrations.

Experiment 1

A 50 cm^3 burette was filled with potassium manganate(VII) solution to the 0.0 cm^3 mark.
25 cm^3 of solution **R** were measured using a pipette and poured into a conical flask.
Potassium manganate(VII) solution was added to the flask while shaking it until a pale pink color appears.

a. The burette showed the following measurement.



final reading

i. Record the final burette reading.
ii. Calculate the volume of potassium manganate(VII) solution used in cm^3.

Experiment 2
25 cm^3 of solution **T** was measured using a measuring cylinder and experiment 1 was repeated. But this time the burette was not filled to the 0.0 cm^3 mark.

b. The initial and final readings of the burette are shown.

initial reading final reading

i. Record the final reading in cm^3.
ii. Calculate the volume of potassium manganate(VII) solution used in cm^3.

c. Which solution is more concentrated solution **R** or solution **T**? Explain

d. i. If in experiment 2, the student used 50 cm^3 of solution **T**, instead of 25 cm^3, calculate the volume of potassium manganate(VII) needed to reach the end point.
ii. What issue will the student face when titrating 50 cm^3 of solution **T**?

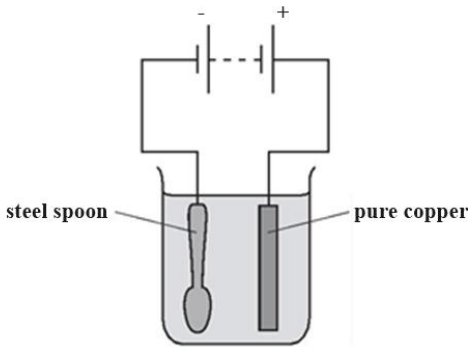
e. Another student conducted the same experiments using a measuring cylinder to measure 25 cm^3 of solutions **R** and **T** instead of a pipette.
How would his results differ from that of the student who used the measuring cylinder?

f. How can the student reach more reliable results in this experiment?

g. Indicators are used in most titrations to determine when the end point is reached. Why no indicator was used in this titration?

Problem Solving Session

1 The apparatus used to electroplate a steel spoon with copper is shown.

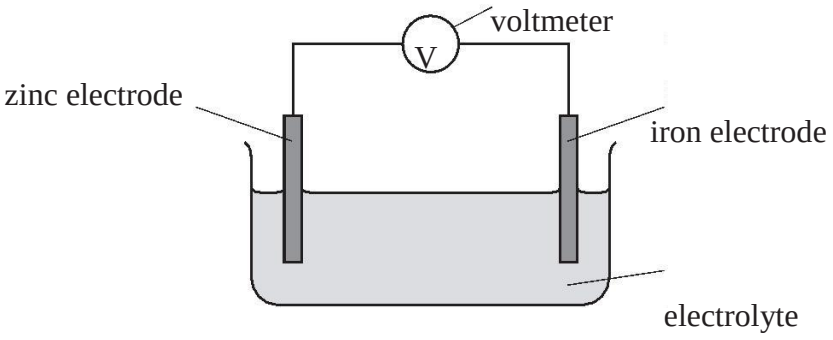


(a) Label the diagram to show:
- the anode
- the electrolyte.

(b) Describe the observations made during this electroplating at the:
steel spoon
copper rod

2 A metal spoon is electroplated with copper.
State what is used as:
the positive electrode (anode) .....
the negative electrode (cathode) .....
the electrolyte ..... [3]
[Total: 3]

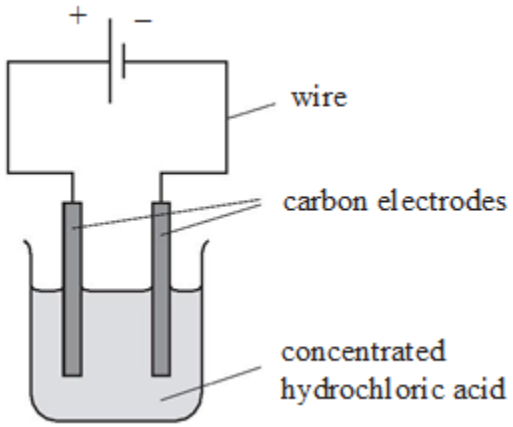
3 This question is about chemical reactions and electricity.
The diagram shows the apparatus used in the production of electrical energy in a simple cell.



The reading on the voltmeter can be increased if either zinc or iron is replaced by another metal.

(a) Name a metal that can replace zinc and increase the reading on the voltmeter.
(b) Name a metal that can replace iron and increase the reading on the voltmeter.

4 This question is about electrolysis.
Concentrated hydrochloric acid is electrolysed using the apparatus shown.



(a) Chloride ions are discharged at the anode.

(i) Complete the ionic half-equation for this reaction.

$$\_\_\_Cl^-(aq) \rightarrow \_\_\_\_(g) + \_\_\_\_ e^-$$

[2]

(ii) State whether oxidation or reduction takes place. Explain your answer.

(b) Describe what is seen at the cathode.

5 This question is about electrolysis.
The table gives information about the electrolysis of two electrolytes.
Carbon (graphite) electrodes are used in each experiment.

(a) Complete the table to show the observations and products of electrolysis.

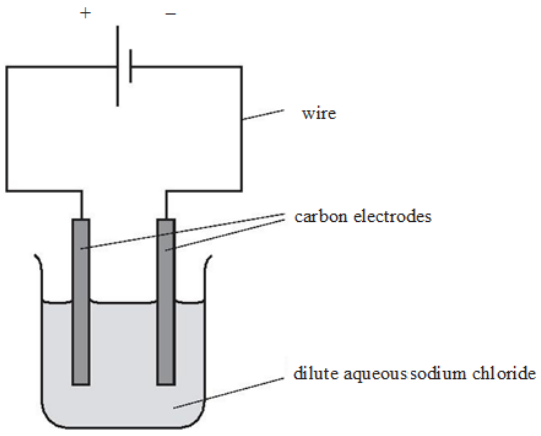
	positive electrode (anode)		negative electrode (cathode)	
electrolyte	observations	name of product	observations	name of product
Aqueous copper(II) sulfate	Colorless bubbles			
concentrated aqueous sodium bromide			Colorless bubbles	hydrogen

(b) Hydrogen is produced at the negative electrode (cathode) during the electrolysis of concentrated aqueous sodium bromide. Write the ionic half-equation for this reaction.

(c) State **two** reasons why carbon (graphite) is suitable to use as an electrode.

(d) Name the particle responsible for the conduction of electricity in the metal wires used in a circuit.

6 A student carries out an electrolysis experiment using the apparatus shown.



The student uses dilute aqueous sodium chloride.
Hydroxide ions are discharged at the anode.

(a) Complete the ionic half-equation for this reaction.

$$OH^-(aq) \rightarrow \_\_\_\_\_ + O_2(g) + 4e^-$$

(b) Explain how the ionic half-equation shows the hydroxide ions are being oxidised.

(c) Describe what the student observes at the cathode.

(d) Write the ionic half-equation for the reaction at the cathode.

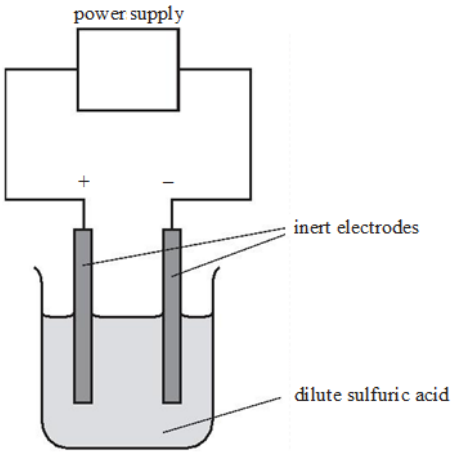
(e) The student repeats the experiment using concentrated aqueous sodium chloride.

(i) Describe what the student observes at:

- the cathode .....
- the anode .....

(ii) The student added litmus to the solution after the electrolysis of concentrated aqueous sodium chloride.
State the colour seen in the solution. Give a reason for your answer.
colour of solution
reason

7 Dilute sulfuric acid is electrolysed using the apparatus shown in the diagram.

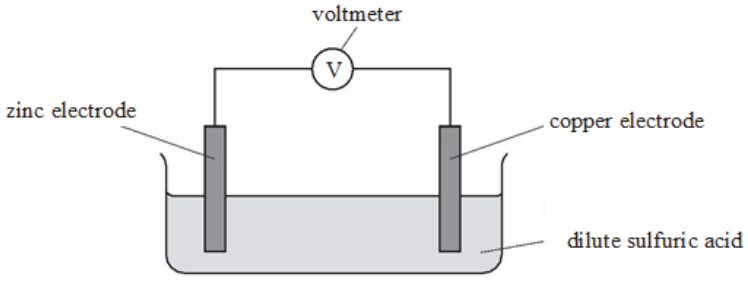


(a) State what is meant by the term *electrolysis*.
(b) Explain why inert electrodes are used.
(c) Name the products formed at each electrode.
(d) Write an ionic half-equation for the reaction at the negative electrode.

8 Electrolysis of concentrated aqueous sodium chloride using inert electrodes forms chlorine.

(a) Name a substance that can be used as the inert electrodes.
(b) Write an ionic half-equation for the formation of hydrogen during this electrolysis.
(c) Give the formulae of the four ions present in concentrated aqueous sodium chloride.
(d) Explain how sodium hydroxide is formed during this electrolysis.

9 A student used the following electrochemical cell.
The reading on the voltmeter was +1.10 V.

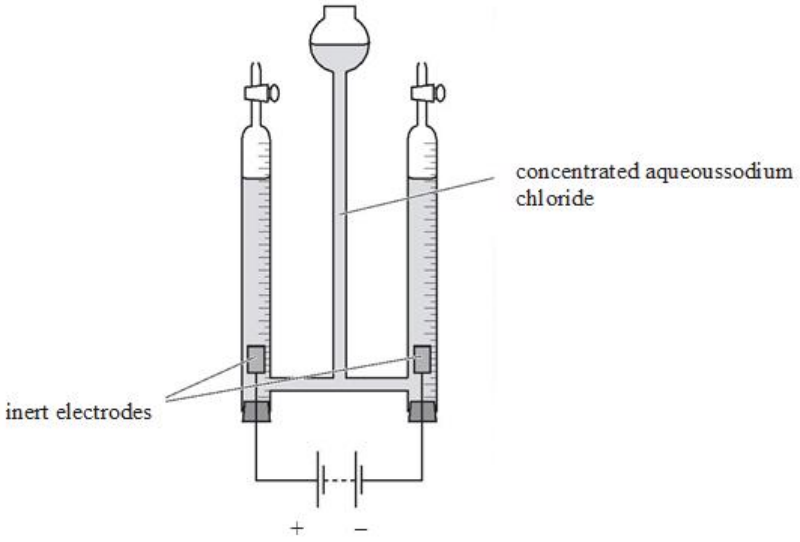


(a) Draw an arrow on the diagram to show the direction of electron flow.
(b) Suggest the change, if any, in the voltmeter reading if the zinc electrode was replaced with an iron electrode. Explain your answer.
(c) The zinc electrode was replaced with a silver electrode. The reading on the voltmeter was –0.46 V. Suggest why the sign of the voltmeter reading became negative.

10 A sample of concentrated hydrobromic acid, HBr(aq), was electrolysed using platinum electrodes. The concentration of the hydrobromic acid was $8.89\,\text{mol}/\text{dm}^3$.

(a) Calculate the concentration of the HBr(aq) in g/dm^3.
(b) Explain why concentrated HBr(aq) can conduct electricity.
(c) Magnesium is **not** a suitable material from which to make the electrodes.
Explain why.
(d) Predict the product formed at the anode when concentrated HBr(aq) is electrolysed.
(e) Write the ionic half-equation for the reaction occurring at the cathode.

11 A student used the following apparatus to electrolyse concentrated aqueous sodium chloride using inert electrodes.



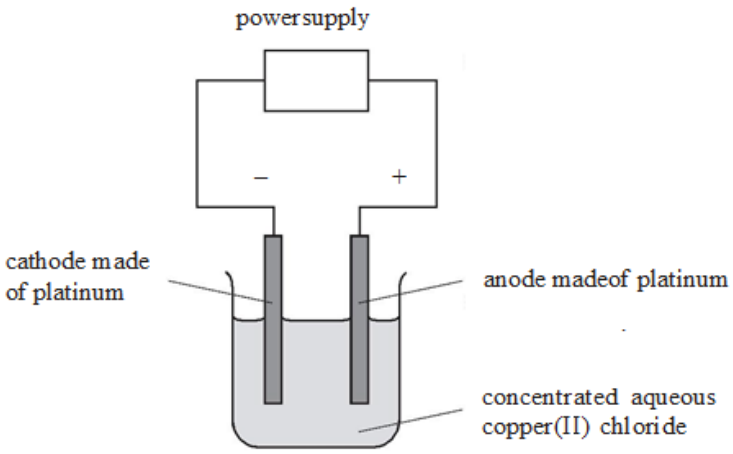
(a) Suggest the name of a metal which could be used as the inert electrodes.
(b) Name the gas formed at the positive electrode.
(c) Write an ionic half-equation for the reaction occurring at the negative electrode.
Include state symbols.
(d) How, if at all, does the pH of the solution change during the electrolysis?
Explain your answer.

12 Metal objects can be electroplated with silver.

(a) Describe how a metal spoon can be electroplated with silver.
Include:
●ù what to use as the positive electrode and as the negative electrode
●ù what to use as the electrolyte
●ù an ionic half-equation to show the formation of silver.
You may include a diagram in your answer.

(b) Give **one** reason why metal spoons are electroplated with silver.

13 Solutions of ionic compounds can be broken down by electrolysis.
Concentrated aqueous copper(II) chloride was electrolysed using the apparatus shown.



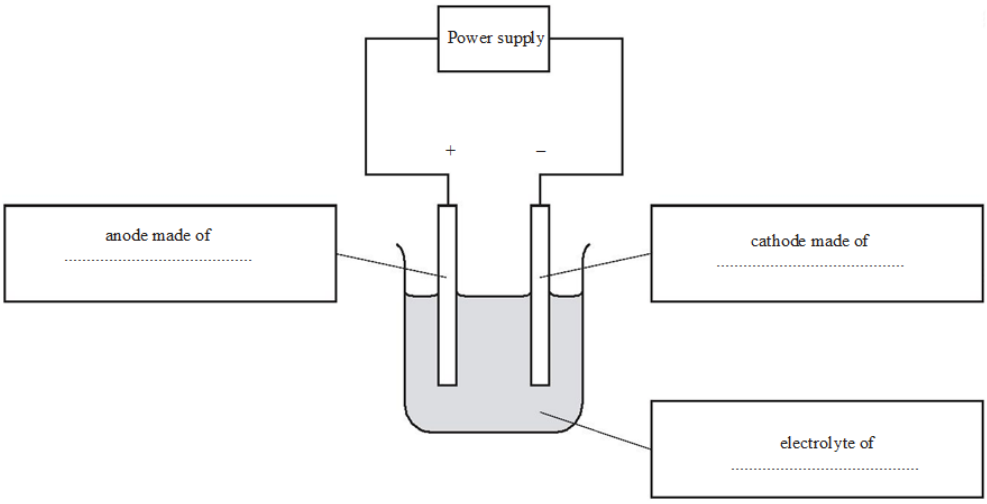
The ionic half-equations for the reactions at the electrodes are shown.

negative electrode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
positive electrode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

(a) Platinum is a solid which is a good conductor of electricity.
State **one** other property of platinum which makes it suitable for use as electrodes.
(b) State what would be **seen** at the positive electrode during this electrolysis.
(c) State and explain what would happen to the mass of the negative electrode during this electrolysis.
(d) The concentrated aqueous copper(II) chloride electrolyte is green.
Suggest what would happen to the colour of the electrolyte during this electrolysis.Explain your answer.
(e) Identify the species that is oxidised during this electrolysis. Explain your answer.

14 Copper is refined (purified) by electrolysis. Nickel can be refined using a similar method.

(a) The diagram shows the refining of nickel by electrolysis.
Complete the labels in the boxes.



(b) Indicate, by writing **N** on the diagram, where nickel is produced.

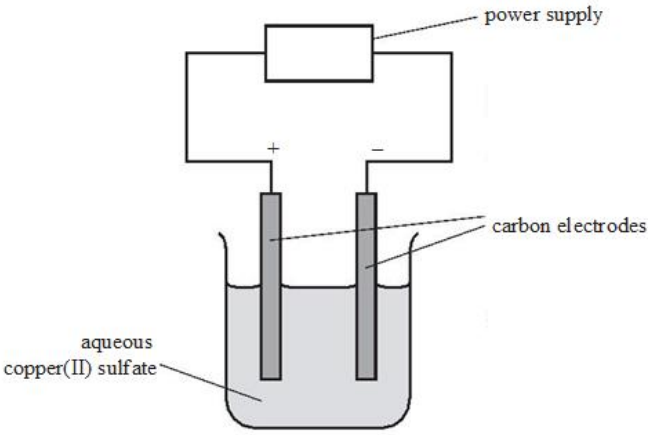
15 The table gives information about the products of the electrolysis of two electrolytes. Platinum
electrodes are used in each case.

(a) Give **two** reasons why platinum is suitable to use as an electrode.

(b) Complete the table.

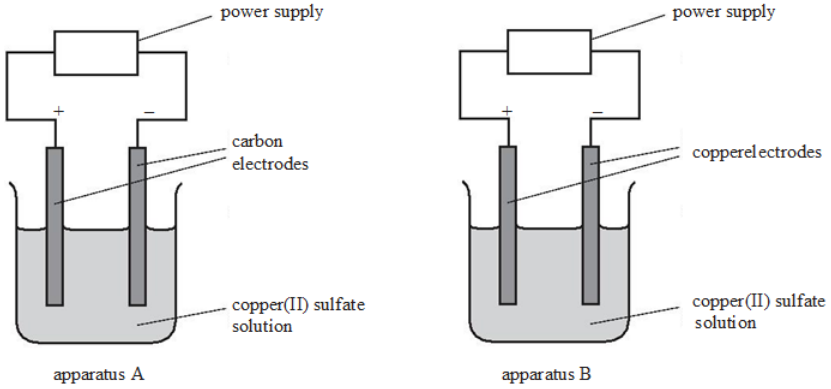
	positive electrode (anode)		negative electrode (cathode)	
electrolyte	observations	name of product	observations	name of product
concentrated aqueous potassium chloride			Colorless bubbles	
Aqueous copper(II) sulfate	Colorless bubbles			

16. A student electrolyses aqueous copper(II) sulfate using the apparatus shown.



Oxygen gas forms at the positive electrode (anode).
What difference would the student observe at the positive electrode if the aqueous copper(II) sulfate were replaced by concentrated aqueous copper(II) chloride?

17 A student electrolysed copper(II) sulfate solution using the two sets of apparatus shown.



In apparatus **A** the student used carbon electrodes. In apparatus **B** the student used copper electrodes. The student made the following observations.

apparatus **A**	apparatus **B**
The mass of the negative electrode increased.	The mass of the negative electrode increased.
The mass of the positive electrode stayed the same.	The mass of the positive electrode decreased.
Bubbles were seen at the positive electrode.	No bubbles were seen at the positive electrode.

Suggest what happens to the colour of the solution in apparatus **A** and apparatus **B** as the electrolysis progresses. Explain your answer.

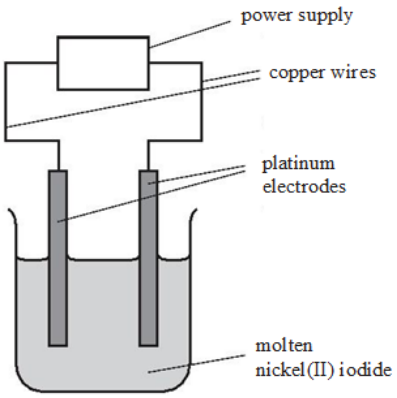
18 Concentrated aqueous potassium bromide is an electrolyte.

(a) Describe the electrolysis of concentrated aqueous potassium bromide.
Include:

- an ionic half-equation for the reaction at the cathode
- the name of the product at the anode
- the name of the potassium compound formed.

(b) When molten potassium bromide is electrolysed, the product at the cathode is different.
Name the product at the cathode when molten potassium bromide is electrolysed.

19 Molten nickel(II) iodide can be electrolysed using the apparatus shown.



During electrolysis, charge is transferred through the copper wires and through the molten nickel(II) iodide.

(a) Name the type of particles which transfer charge through the copper wires.
(b) Name the type of particles which transfer charge through the molten nickel(II) iodide.
(c) Predict the products of the electrolysis of molten nickel(II) iodide. Write an ionic half-equation for the formation of **one** of these products.

Chapter 5 - Stoichiometry

1. **At RTP** (25℃ , 1 atm), 1 mole of a gas occupies **24 dm^3** volume
2. **At STP** (0℃ , 1 atm), 1 mole of a gas occupies **22.4 dm^3** volume
3. Stoichiometric exercises.

CQD 1. [G]
a. When 4.00 mol of Fe react with excess chlorine gas, what mass of $FeCl_3$ will be produced?
[Fe = 56.0; Cl = 35.5]
b. What volume of chlorine at STP reacts with 4.00 moles of Fe?

CQD 3. [G] Hydrogen gas, H_2, reacts with carbon monoxide gas, CO, to produce methanol, CH_3OH, as shown in the following reaction.

$$2H_2 + CO \rightarrow CH_3OH$$

When 0.50 tons of H_2 gas reacts completely with CO gas, the amount of methanol produced was 3.0 tons. Calculate the percentage yield. [H = 1.00; C = 12.0; O = 16.0]

CQD 17. The metal zinc, Zn, reacts with acid to release hydrogen gas, H_2.
What volume of hydrogen gas would be produced at room temperature and pressure when 3.27 g of Zn react completely with excess sulfuric acid? [Zn = 65], $V_M = 24\ dm^3\ mol^{-1}$

4. Find empirical formula (simplest formula) of an organic compound given its percentage composition

CQD 20. [G] A hydrocarbon is 75 percent by mass carbon. Find its empirical (simplest) formula.
[C = 12.0; H = 1.00]

5. Determining molecular formula of a compound given empirical formula and molar mass of the molecular formula

CQD 19. [G] The empirical formula of a compound is CH_2O and its molar mass is 60 g/mol. Find the molecular formula of the compound.
[C = 12.0; H = 1.00; O = 16.0]

Ex. 10 The empirical formula of a compound is C_2H_7N. If the molar mass of the compound is 135 g/mole. Find the molecular formula.

CQD 21. [T]
During a lab session, students were asked to prepare hydrated magnesium sulfate by adding an excess of dilute sulfuric acid to magnesium oxide.

Each student was handed a different mass of magnesium oxide.

$$\underset{M_f = 40.3}{MgO} \xrightarrow{\text{excess } H_2SO_4} \underset{M_f = 120.4}{MgSO_4.7H_2O}$$

At the end of the experiment, each student dried and weighed the crystals obtained.
Which student produced the highest percentage yield of hydrated magnesium sulfate?

	mass of magnesium oxide used / g	mass of crystals produced / g
A	5.0	9.5
B	10.0	24.0
C	16.0	35.8
D	24.0	46.6
E	36.0	56.7

CQD 22. [T]Ammonium sulfate, $(NH_4)_2SO_4$ is used as fertilizer.
It is produced from ammonia and sulfuric acid as shown in the chemical equation.

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

Calculate the maximum mass of ammonium sulfate produced from 680 g of ammonia, NH_3, using the following steps. The relative formula mass of ammonia M_r = 17.

a) Calculate the moles of NH_3 in 680 g.
b) Deduce the number of moles of $(NH_4)_2SO_4$.
c) Calculate the M_r of $(NH_4)_2SO_4$.
d) Calculate the maximum mass of ammonium sulfate produced.

CQD 23. [T]
a. The relative atomic mass of an atom is given by comparing the mass of that atom against a standard atom.

i. What is the symbol of the standard atom? Include its mass number.
ii. Ammonia has relative molecular mass, M_r of 17 and lithium fluoride has a relative formula mass, M_r of 26.
A student claims that the term 'relative molecular mass' can be used for both compounds, ammonia and lithium fluoride.
Do you agree with him? Explain your reasoning.

b. A 0.174 g of gas **X** has a volume of 72 cm^3 at room temperature and pressure.

i. How many moles of gas **X** in 72 cm^3? V_M = 24 dm^3 mol^{-1}.
ii. What is the relative molecular mass of gas **X**?
iii. It is found that gas **X** is a hydrocarbon of empirical formula C_2H_5.
Find the chemical formula of gas **X**. Show your work. [C = 12.0; H = 1.00]

c. A 12 g of an oxide of iron is produced from burning 8.4 g of iron.
Calculate the empirical of this oxide of iron. [Fe = 56.0; O = 16.0]

d. Another oxide of iron has the empirical formula of Fe_3O_4.
One molecule of this oxide of iron contains six atoms of iron.
Deduce the molecular formula of the oxide.

Problem Solving

1 When 80 g of iron(III) oxide reacts with excess carbon monoxide, 56 g of iron is produced.
Calculate the minimum mass of iron(III) oxide needed to produce 14 g of iron.

2 Dilute sulfuric acid and aqueous potassium hydroxide are used to make aqueous potassium sulfate.

$$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$$

Calculate the volume of 0.0625 mol / dm^3 dilute sulfuric acid, H_2SO_4, that completely reacts with 25.0 cm^3 of 0.100 mol / dm^3 potassium hydroxide, KOH, to produce aqueous potassium sulfate.

Use the following steps.

(a) Calculate the number of moles of KOH in 25.0 cm^3 of 0.100 mol / dm^3 KOH.

(b) Deduce the number of moles of H_2SO_4 that react with KOH.

(c) Calculate the volume of H_2SO_4 required.

(d) The experiment is repeated using the same volume and concentration of potassium hydroxide and the same concentration of dilute sulfuric acid. In this second experiment, the product is aqueous potassium hydrogensulfate, $KHSO_4$.

$$H_2SO_4 + KOH \rightarrow KHSO_4 + H_2O$$

Use your answer to **(c)** and the equation to deduce the volume of H_2SO_4 required.

3 Calcium carbonate decomposes when heated.

$$\text{calcium carbonate} \rightarrow \text{calcium oxide} + \text{carbon dioxide}$$

When 0.50 g of calcium carbonate decomposes, 120 cm^3 of carbon dioxide gas is produced.
Calculate the volume of carbon dioxide gas produced when 0.10 g of calcium carbonate is used.

4 Lead is a metallic element in Group IV. One of the ores of lead is galena, which is an impure form of lead(II) sulfide, PbS.
Lead also occurs in the ore cerussite, which contains lead(II) carbonate, $PbCO_3$.

(a) Calculate the relative formula mass, M_r, of $PbCO_3$.

(b) The M_r of PbS is 239.
Calculate the percentage of lead by mass in PbS.

5 In an experiment, 1.61 g of $Na_2SO_4 \bullet \boldsymbol{x}H_2O$ is heated until all the water is given off. The mass of Na_2SO_4 remaining is 0.71 g.
[M_r: Na_2SO_4,142; H_2O,18]

Determine the value of ***x*** using the following steps.

- Calculate the number of moles of Na_2SO_4 remaining.
- Calculate the mass of H_2O given off.
- Calculate the number of moles of H_2O given off.
- Determine the value of ***x***

6 One mole of ethane, C_2H_6, contains 6.02×10^{23} molecules.

Calculate how many molecules are in 15g of ethane.

7 On analysing hydrated iron(II) sulfate crystals, $FeSO_4 \bullet xH_2O$, a student found that one mole of the crystals had a mass of 278 g.

Determine the value of *x* using the following steps:

- calculate the mass of one mole of $FeSO_4$
- calculate the mass of H_2O present in one mole of $FeSO_4 \bullet xH_2O$
- determine the value of *x*.

8 Ammonia nitrate, NH_4NO_3, is a fertiliser.
Calculate the percentage by mass of nitrogen in ammonium nitrate.

9 Part of the definition of relative mass is 'the average mass of naturally occurring atoms of an element'.

Some relative atomic masses are not whole numbers.

Elements **Y** has only two different types of atoms, ^{69}**Y** and ^{71}**Y**.
The ratio of atoms present in element **Y** is shown.

$$^{69}\mathbf{Y} : {}^{71}\mathbf{Y} = 3 : 2$$

- Calculate the relative atomic mass of element **Y** to **one decimal place**.
- Identify element **Y**.

10 Tin is a metallic element in Group IV. Its main ore is cassiterite which is an impure form of tin(IV) oxide, SnO_2.
Tin also occurs in stannite, Cu_2FeSnS_4.

(a) Calculate the relative formula mass, M_r, of Cu_2FeSnS_4.

(b) The M_r of SnO_2 is 151.
Calculate the percentage of tin by mass in SnO_2.

(c) The percentage of tin by mass in Cu_2FeSnS_4 is 27.6 %.
Use this information and your answer to **(b)** to suggest whether it would be better to extract tin from SnO_2 or Cu_2FeSnS_4.Explain your answer.

11. In a titration, 25.0 cm^3 of 0.0800 mol / dm^3 aqueous potassium hydroxide, KOH(aq), is neutralised by 20.0 cm^3 of dilute sulfuric acid, H_2SO_4(aq).

$$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

Calculate the concentration of H_2SO_4, in g / dm3 using the following steps.

- Calculate the number of moles of KOH used.
- Determine the number of moles of H_2SO_4 which react with the KOH.
- Calculate the concentration of H_2SO_4 in mol / dm3.
- Calculate the concentration of H_2SO_4 in g / dm3.

12. Which sample contains the largest number of molecules?

A 16 g of methane, CH_4(g)
B 16 g of oxygen, O_2(g)
C 16 g of phosphorus, P_4(s)
D 16 dm^3 of methane at r.t.p., CH_4(g)

Chapter 7-Electrons and The Periodic Table

1. Recall the three subatomic particles, their relative masses and charges.
2. Recall that atomic or proton number is the number of protons in an atom.
3. Recall that mass or nucleon number is the number of protons and neutrons in an atom.
4. Recall that isotopes are atoms of the same element having the same atomic number but different mass number.
5. Recall how to find the average atomic mass using isotopic masses and % abundance.

Ex1. The element copper has two naturally occurring isotopes of mass numbers of 63 and 65. The relative abundances are 69.2% and 30.8% respectively. Calculate the average atomic mass of copper.

6. An atom consists of a central dense region, called a nucleus where protons and neutrons (collectively called nucleons) are found, surrounded by electrons.
7. The areas around the nucleus where electrons have a high probability of being found are defined by shell numbers (or energy levels). As the shell's distance away from the nucleus increases, its energy increases.
8. Electronic configurations can be written in many ways, simplest way is using the number of electrons occupying the energy levels. The first energy level can hold a maximum of 2 electrons, the second a maximum of 8, the third a maximum of 18 and so on. Therefore, the electronic configuration of the first twenty elements is as shown below:

$_{1}$H: 1	$_{2}$He: 2
$_{3}$Li: 2,1	$_{4}$Be: 2,2
$_{5}$B: 2,3	$_{6}$C: 2,4
$_{7}$N: 2,5	$_{8}$O: 2,6
$_{9}$F: 2,7	$_{10}$Ne: 2,8
$_{11}$Na: 2,8,1	$_{12}$Mg: 2,8,2
$_{13}$Al: 2,8,3	$_{14}$Si: 2,8,4
$_{15}$P: 2,8,5	$_{16}$S: 2,8,6
$_{17}$Cl: 2,8,7	$_{18}$Ar: 2,8,8
$_{19}$K: 2,8,8.1	$_{20}$Ca: 2,8,8,2

CQD 4. Write the electronic configuration of the following atoms: $_{11}$Na, $_{15}$P, $_{19}$K

CQD 7. [G] Write the electron configurations of the following atoms: $_{3}$Li, $_{15}$P.

CQD 8. [G] Write the electron configuration of the following atoms using 2, 8, 8 electronic configurations: $_{11}$Na, $_{8}$O.

9. The Periodic Table is arranged in Periods (horizontal rows), groups (vertical rows) and blocks (rectangular areas of the Periodic Table**).**
10. **Periods:** are the horizontal rows of the Periodic Table.
11. The period in which an element is found, can be predicted from the number of shells it has. Example:

- Hydrogen and helium are the only elements in period 1 because they have electrons in the first shell only ($_{1}$H: **1**, $_{2}$He: **2**)

- An element whose electronic configuration is 2, 8,1 belongs to period 3 as its three shells.

12. **Groups:** are the vertical rows of the Periodic Table**.**

13. The group in which an element is found, can be predicted from its number of valence electrons. Also remember that all members of the same group have the same number of valence electrons.

14. Valence electrons: are the electrons present in the last shell.

An element whose electronic configuration is (2,8,**4**), has 4 valence electrons so it belongs to group 4.

15. **Core electrons:** are the inner shell electrons that are strongly bound to the nucleus.

Example: Nitrogen, $_{7}N$

a) The electronic configuration is 2,5
   - Number of valence electrons: 5
   - Number of core electrons: 2

16. The electronic configuration of noble gases ends in 8, this gives the Noble gases their special inertness (unreactive) as the outer shell is full. They have no tendency to combine with any other atom and so they exist as monoatomic gases. As a matter of fact, other elements tend to gain, lose or share electrons to achieve the same electronic structure as the Noble gases.

- **Groups 1, 2 and 3 Elements:** react by losing electrons to reach electronic configuration of the nearest noble gas. For example: $_{12}Mg$, has 2 electrons more than neon, $_{10}Ne$, so it reacts by losing $2e^{-}$'s.
- **Groups 5, 6 and 7 Elements:** react by sharing or gaining 3, 2 or 1 electron respectively to reach the highest electronic configuration of the nearest noble gas. For example: $_{8}O$, has 2 electrons less than neon, $_{10}$ Ne, so it reacts by gaining or sharing $2e^{-}$'s.
- **Group 4 Elements:** neither gain nor lose electrons. They tend to share electrons to reach the electronic configuration of the nearest noble gas.

17. Isotopes of the same element have the same chemical properties because they have the same number of electrons and therefore the same electronic configuration. Some isotopes show an imbalance in their nucleus making them unstable and tend to spontaneously break up emitting certain radiations. Such isotopes are called radioactive. Examples include carbon-14 and tritium.

18. Some uses of radioactive isotopes include:

- energy source
- as tracers in industries and hospitals to trace the flow of substances through a system. Example: doctors use radioactive iodine to find out if a patient's kidney is blocked.
- dating, used to find out how old a material is (carbon and uranium isotopes).

19. Elements in the same group of the Periodic Table show similarities in their chemical properties because their electronic configurations all end with the same number (number of valence electrons). This is why the position of an element in the Periodic Table can be used to predict its properties.

Ex2.

a. What is special about the electronic configuration of the noble (inert) gases?
b. Is this, the electronic configuration of a noble gas: 2, 8, 7?

Ex3. Fill the table below:

Element	number of electrons in each energy level	Period	# of valence e^-	Group
$_{11}Na$	2, 8, 1	3	1	1
$_{3}Li$				
$_{19}K$				
$_{15}P$				
$_{18}Ar$				

20. **Ions:** are charged species formed by the loss or gain of one or more electron**.**

Positive Ions (Cations)	**Negative Ions (Anions)**
Formed by the loss of one or more electron	Formed by the gain of one or more electrons
$X \xrightarrow{-1\,e^-} X^{+1}$	$Y \xrightarrow{+1\,e^-} Y^{-1}$
The number of protons is greater than the number of electrons	The number of electrons is greater than the number of protons
Metals tend to form positive ions	Non-metals tend to form negative ions

21. Electronic configuration of ions can be written only after determining the number of electrons in the ion. To do that for a

- CATION: consider the atomic number of the atom, subtract from it the cation's charge.
- ANION: consider the atomic number of the atom, add to it the anion's charge.

For example:
$_{13}Al^{+3}$ the atom has 13 e^-s, the ion will have 13 – 3 = 10 electrons. Its configuration will be: **2,8**
$_{9}F^{-1}$ the atom has 9 e^-s; the ion will have 9 + 1 = 10 electrons. Its configuration will be: **2,8**

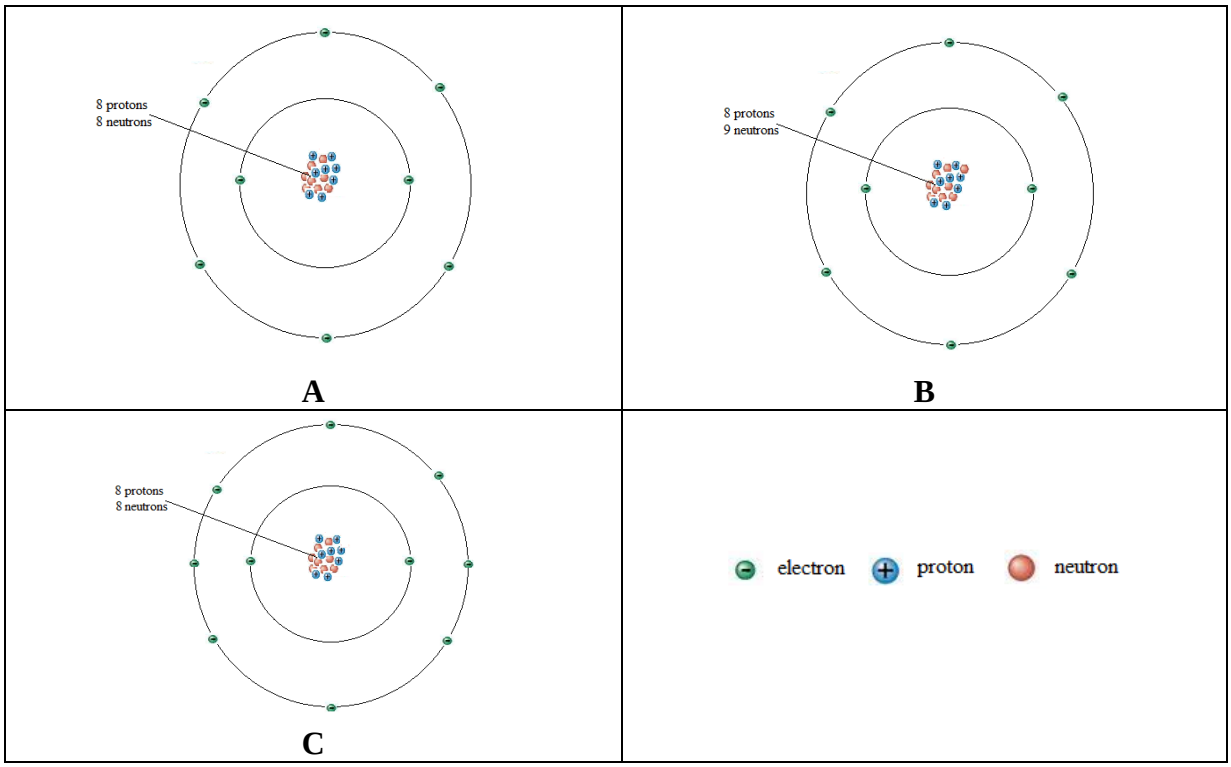
CQD5. Write the electronic configuration of the following ions:

a) $_{8}O^{-2}$
b) $_{11}Na^{+}$
c) $_{12}Mg^{+2}$
d) $_{7}N^{-3}$

CQD6. Use the Periodic Table to help write the formula of the following compounds or ions

Compound or ion	Formula	Compound or ion	Formula
Barium Oxide		Barium chloride	
Boron oxide		Aluminum nitride	
Aluminum fluoride		Aluminum ion	
Sodium sulfide		Sulfide ion	
Calcium nitride		Nitride ion	
Potassium oxide		Phosphide ion	
Chloride ion		Calcium ion	

CQD15. [T] The table below shows Bohr's model for three particles **A**, **B** and **C**.



a. Give the reason why

i. particle **C** is a negative ion.
ii. **A**, **B** and **C** are particles of the same element?
iii. **A** and **B** are isotopes of the same element.
iv. **A** is a nonmetal.

b. i. Use the periodic table to identify **A**. Write its atomic symbol.
ii. Write the atomic symbol for particle **C**.

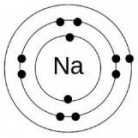
c. An atom **X** has 16 electrons.
What is the electronic structure of **X**?

Chapter 8 – Molecules in Gas Phase

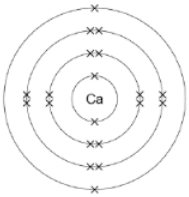
1. **Chemical bond:** is the force that holds atoms together and makes them function as one unit.
2. **Bond length:** is the distance between the nuclei of two bonded atoms.
3. **Covalent bond**: is the equal sharing of one or more pairs of e⁻s between 2 non-metal atoms leading to a noble gas electronic configuration.
4. **Atoms may form a:**
   a) **Single covalent bond**: one pair of electrons is shared.
   b) **Double covalent bond**: 2 pairs of electrons are shared.
   c) **Triple covalent bond**: 3 pairs of electrons are shared.
5. **Ionic bond:** is the electrostatic force of attraction between oppositely charged ions.
6. **Cation:** a positive ion. **A** metal atom tends to lose one or more electrons to attain the electronic configuration of the noble gas before it thus forming a positive ion.
7. **Anion:** a negative ion**. A** non-metal atom tends to gain one or more electrons to attain the electronic configuration of the noble gas after it thus forming a negative ion.
8. **Electron dot representation of atoms:** consists of the **chemical symbol** for the element surrounded by dots representing its electrons. The chemical symbol represents the atom's **nucleus**. We may draw all the electrons or only the valence ones depending on what is required

Ex1. Draw the dot cross diagrams of the following atoms:

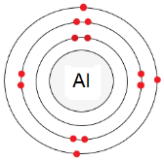
a) Na



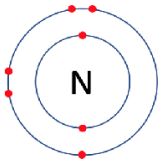
b) Ca



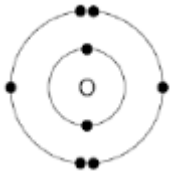
c) Al



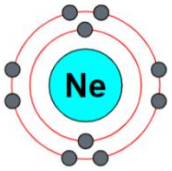
d) N



e) O

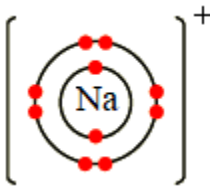


f) Ne



9. Electron dot representation of cations

Example, Na^+:

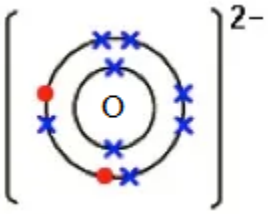


Ex2. Draw the electron dot cross diagram of the following species:

a) Mg^{2+}
b) Al^{3+}

10. Electron dot representation of anions: The gained electrons are paired with the original valence electrons and represented in a different symbol.

For eg., O^{2-}:

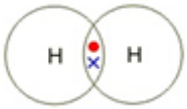


Ex3. Draw the electrons dot cross diagrams of the following species

a) S^{2-}
b) N^{3-}
c) Cl^-

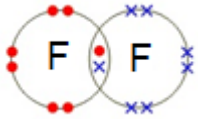
11. Electron dot cross diagrams of elements

Example: Hydrogen, H_2

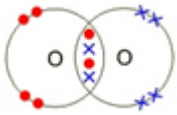


Q6 Draw the electron dot representations of each of the following:

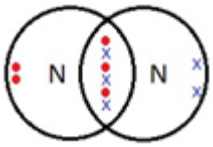
a) Fluorine, F_2



b) Oxygen, O_2

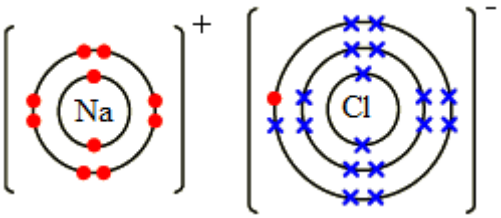


c) Nitrogen, N_2

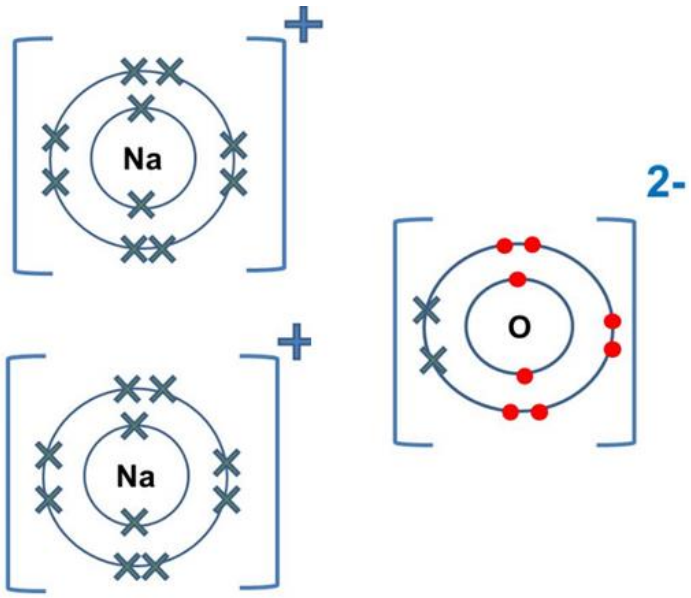


12. Electron dot cross diagram of ionic compounds: To draw the diagram:

a) Decide on the atoms required, the number of each and the electronic structure.
b) Assume complete transfer of electrons between atoms to form ions with noble gas electronic structure.
c) Show only the outer energy level electronic structure. Draw square brackets around each ion and show its charge.
For eg.;
Sodium chloride, NaCl



Sodium oxide, Na_2O

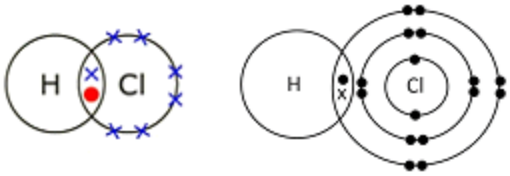


Ex4. Draw the electrons dot cross diagrams of the following compounds:

a. CaO
b. Al_2O_3

13. Electron dot cross diagrams of covalent compounds

Example: Hydrogen chloride, HCl



Ex5. Draw the electrons dot cross diagrams of the following compounds:

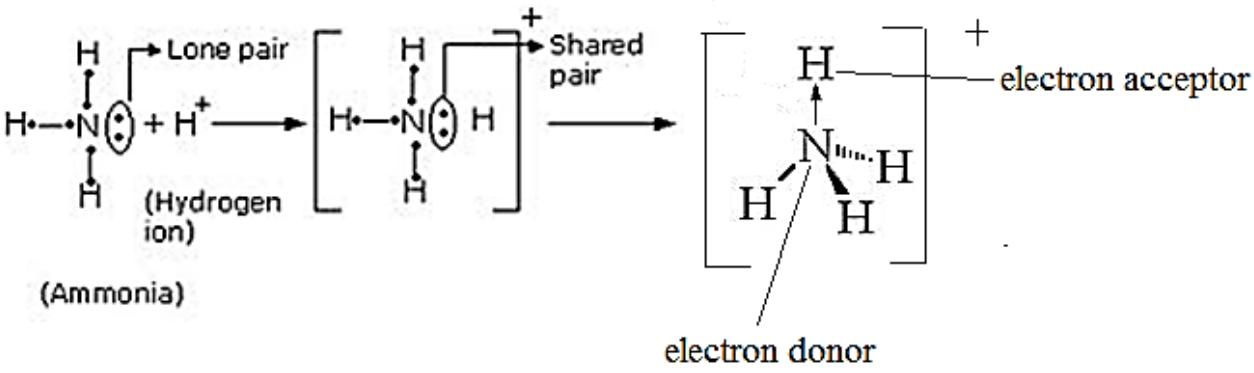
a) H_2O
b) AsH_3
c) CH_4
d) SiH_4
e) CO_2

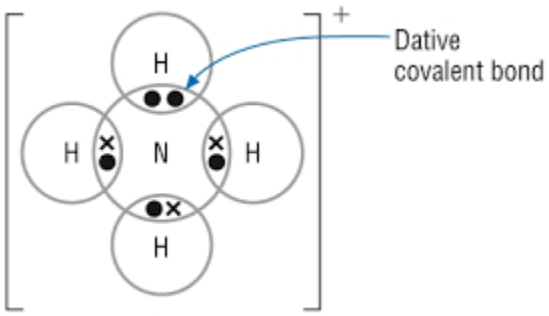
CQD2. [G] Give the electron dot diagram for the each of the following chemical species.

a. H_2
b. HF
c. H_2S
d. NH_3
e. CH_4
f. NaI
g. $MgCl_2$
h. Cl_2
i. CH_3OH

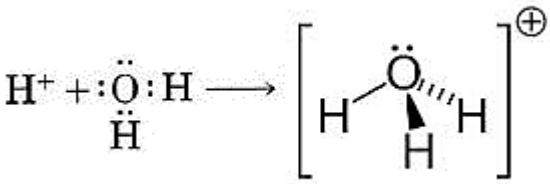
14. Dative (coordinate) covalent bond: is a covalent bond formed when both electrons in the bond come from the same atom**.** Example of species with dative covalent bond include:

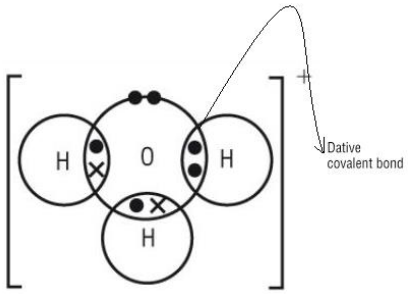
Ammonium ion, NH_4^+, $NH_3 + H^+ \rightarrow NH_4^+$



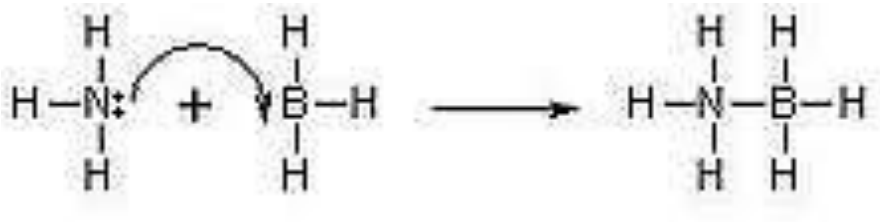


Hydronium ion, H_3O^+, $H_2O + H^+ \rightarrow H_3O^+$





NH_3BF_3, $NH_3 + BF_3 \rightarrow NH_3BF_3$

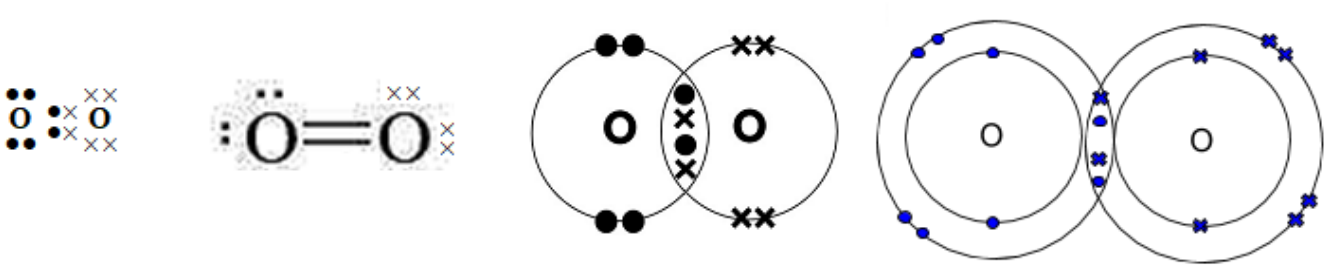


CQD4. [G] Define a dative (coordinate) covalent bond.

Ex6. Which of the following molecules/ ions possess coordinate covalent bonding?

BF_4^-, CO_3^{2-}, H_3O^+, NH_3, NH_4^+

15. The Oxygen molecule, O_2.



CQD6. Draw the cross-dot diagram for O_2, C_2H_4 and N_2.

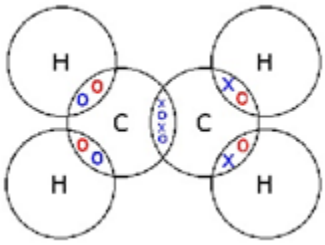
CQD7. a. Draw the electron dot diagram for CO_2 molecule.

b. What is the shape of the CO_2 molecule?

c. Does it have polar bonds?

d. Is it a polar molecule?

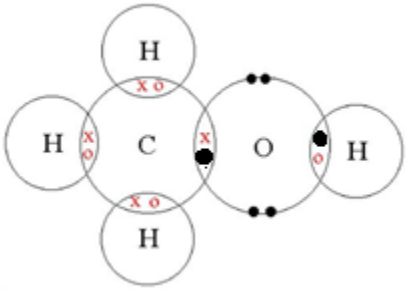
16. Structure of the ethene molecule, C_2H_4



Q18 Which of the following is/are **<u>true</u>** regarding the C_2H_4 molecule?

a) It is a linear molecule
b) It is non-polar
c) The C atoms use sp^3 hybrid orbitals
d) The C and H atoms are covalently bonded
e) The C and C are double bonded
f) It is a triangular molecule
g) It is a planar molecule

17. Methanol, CH_3OH



CQD8. [T]

Arrangement and movement of particles in solid, liquid or gas are explained by the kinetic theory.

a. At room temperature, chlorine exists in the gas state.
i. Give the Lewis dot structure of a chlorine molecule.
ii. Compare the movement and arrangement of chlorine molecules in the solid state to those in the gas state.
iii. The pressure inside a sealed container containing chlorine gas increases with temperature. Use the kinetic molecular theory to explain the relation between the pressure of a gas and its temperature.

b. The following table provides the rate of diffusion of chlorine and hydrogen chloride at 25°C.

Gas	Rate of diffusion in cm^3/min
Chlorine	0.63
Hydrogen chloride	0.87

i. Why hydrogen chloride diffuses faster than chlorine?
ii. If temperature is increased both gases diffuse faster. Why?
iii. Compare the rate of diffusion of carbon dioxide gas to that of chlorine and that of hydrogen chloride.

CQD9. (T) a. Draw the dot-and-cross diagram to show the electron arrangement in one molecule of methanol. Show the outer shell electrons only.
b. Complete the dot-and-cross diagram to show the electron arrangement in a molecule of NF_3. Use dots for nitrogen electrons and crosses for fluorine electrons. Show outer electrons only.

Chapter 9-Bonding in Solids and Liquids

1.

Types of Bonding	
Intermolecular forces of attraction	**Intramolecular forces of attraction**
Occurs between molecules	Occurs within the molecule
Represented as (- - - -)	Represented as (——)
Van der Waal Forces	Covalent bond (between 2 non-metals)

2. Summary of bonds in a crystal:

Element or compound	**Type of intermolecular forces**
Metals	Metallic bond
Ionic compound	Ionic bond (between metal and non-metal)
Non-metal Elements: a) monatomic as He, ne, Kr b) polyatomic as S_8, P_4Cl_2, O_2, N_2, H_2	VDW
Carbon (in diamond or graphite) or Si (in SiO_2, sand or silicon dioxide)	Covalent bond
Molecular compound	VDW

3. Types of Crystalline Solids

I. Ionic solids: (Ions occupy lattice points forming giant strucrures, example sodium chloride)
II. Molecular solids: (Discrete covalent molecules occupy lattice points. Example: ice and sucrose)
III. Atomic Solids (pure elements)
IV. Metallic solids (form giant structures)
V. Network solids (form giant structures)
VI. Group 8A solids

4. A giant structure is a lattice where the structure repeats itself in all directions, the forces involves are the same in all directions holding the whole structure together.

5. Ionic Crystals:

Definition:

An ionic bond is a strong electrostatic force of attraction between positively charged ions (cations) formed by metals and negatively charged ions (anions) formed by non-metals.

An ionic compound forms a giant ionic lattice structure of regularly arranged, alternating positive and negative ions.

CQD21. [G] Describe the structure of the solid ionic compound NaBr.

Properties:

1) Hard crystalline solids (*regular arrangement of ions in a lattice with the ions of opposite charge next to each other*).
2) Have high melting point and boiling point (*due to strong electrostatic force of attraction between oppositely charged ions, large amount of energy is needed to separate them*).
3) Usually soluble in water (*as water molecules are able to bind with positive and negative that make up the lattice*).
   Lattice is a regular pattern of positive and negative ions.
4) Do not conduct electricity when solid but conduct electricity when molten or aqueous.
   When solid, ions are not free to move and carry electric current, but when molten or in aqueous solution, ions are free to move and carry electric current.
5) Brittle (*due to its layer structure, pushing one layer against another brings the ions of same charge next to each other, the repulsive forces resulting force the layers apart*).

Diagram:

Every Na^+ is surrounded by 6 Cl^-
Every Cl^- is surrounded by 6 Na^+

Ex1.

i. Which of the following is/are **correct** properties of ionic compounds?
   b) Ductile
   c) Brittle
   d) Malleable
   e) Lattice structure
   f) Hard
   g) Low melting points
   h) High boiling points
   i) Good conductors of electricity in all states
   j) Form regularly shaped crystals

ii. List the properties possessed by this compound.

CQD18. Explain why solid NaCl is a non-conductor of electricity, while when melted it becomes a good electrolyte.

CQD11. Magnesium reacts with chlorine to form magnesium chloride. In the lattice of magnesium chloride, the ratio of magnesium ions to chloride ions is 1:2. Explain the term 'lattice'.

6. Covalent Crystals:

Definition:

A covalent bond is formed when a pair of electrons is shared between two atoms leading to noble gas electronic configurations.

Properties:

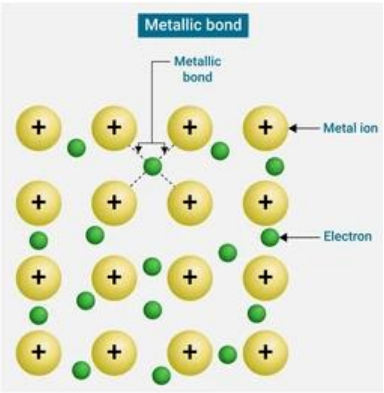
1) They are gases or volatile liquids or low melting point solids(*they are made of simple molecules and the intermolecular forces are very weak so not much energy is needed to move molecules further apart*).
2) Have low melting point or boiling point (*they are made of simple molecules and the intermolecular forces are very weak so not much energy is needed to move molecules further apart*).
3) Insoluble in water but soluble in organic solvents.
4) Do not conduct electricity when molten or in aqueous solutions (*because they do not have free ions or delocalized electrons that can move and carry electric current*).

CQD14. Explain why an aqueous solution of sugar is a non-electrolyte, while that of NaCl is a good electrolyte.

7. Metallic Crystals

Definition:

A metallic bond is an electrostatic force of attraction between the positive ions in a giant metallic lattice and a delocalized sea of electrons.



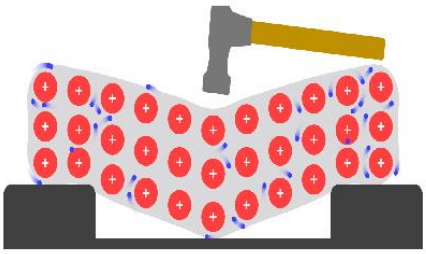
The structure of a typical metal

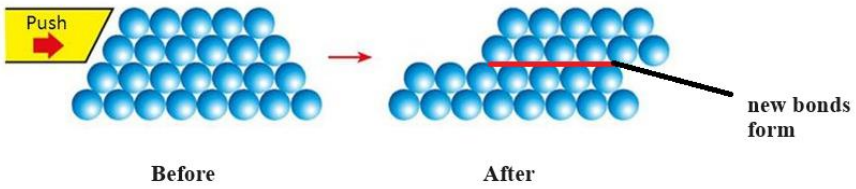
It is formed of positively charged ions arranged in a lattice surrounded by freely moving electrons.

CQD20. [G] Describe metallic bonding.

Properties:

Property	Reason
High melting point and boiling	Strong extensive metallic bonding requires large amounts of energy to be overcome
Conducts heat and electricity	Due to mobile sea of free electrons or delocalized electrons that are free to move and carry electric current.
Malleable and ductile	When a force is applied, the layers of positive ions can easily slide over one another. The attractive forces in metallic bonding act in all directions so when layers slide, new bonds are easily formed.
Shiny and lustrous	Because delocalized electrons can absorb and re-emit light energy which is visible as shine.





Alloys that are metal solutions (such as steel, brass and bronze) share the properties of metals. Due to impurities, they will be harder and stronger than pure metals but of lower conductivity,

CQD4. Explain the following observation:
Sulfur is a non-conductor element while sodium is a conductor of heat and electricity.

CQD5. Explain why silver is a shiny element.

CQD6. Explain why sulfur is brittle while copper is ductile and malleable.

CQD7. [G] Give four general properties of metals.

Ex2.

a. In which of the following substances are the intermolecular forces mainly metallic?
H_2; Na ; O_2 ; C.
b. Where do you generally find metallic elements in the Periodic Table?
c. Is it easy for metals to lose electrons?
d. As a result, can they act as oxidizing agents or as reducing agents?

Ex3. Which of the following substances conducts electricity in the solid state?

a) sodium chloride
b) Diamond
c) Lead (II) chloride
d) Chlorine
e) Potassium
f) Silver nitrate

Ex4.

a. What is meant by metallic bonding?
b. Many of the physical properties of metals can be explained using the idea of _____
c. List some elements that are conductors. What makes them conductors?

8. Network Solids

Giant covalent structures shown by some elements and some compounds.

Giant atomic structure	Giant molecular structure
Basic unit: atom	Basic unit: molecule
Examples Diamond Graphite	Examples: Sand (Quartz or silicon dioxide or silicon (IV) oxide or SiO_2)

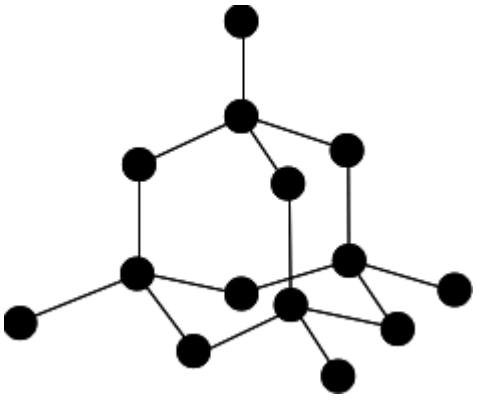
Giant atomic structure

Allotropy

Allotrope means when two (or more) forms of an element exist in the same physical state.
Allotropes of carbon are: diamond, graphite and buckministerfullerene

Diamond:

Strong covalent bonds hold the C-atoms in all directions. Each C-atom is bonded to four others arranged tetrahedrally.



Properties:

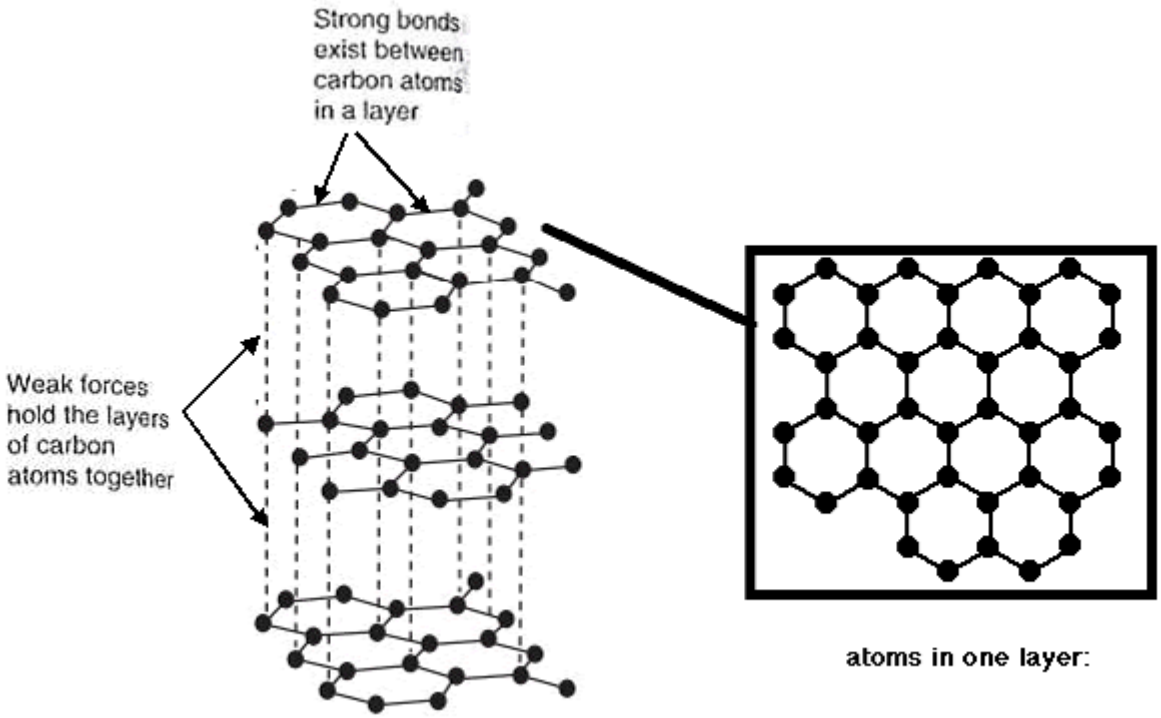
- Very high melting point (*a large number of strong covalent bonds extending throughout the structure need to be broken which requires huge amounts of energy*).
- Very hard (*the bonds are rigid*)
- No electric or thermal conductivity (*no free moving charged particles*)

Graphite:

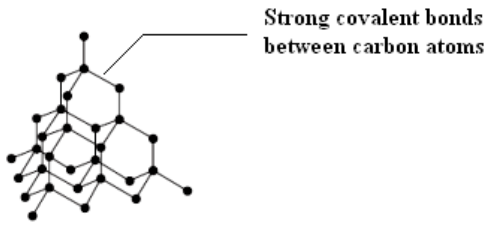
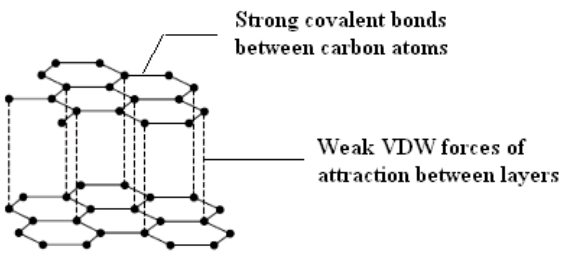
Strong covalent bonds hold the C-atoms two dimensions, each C-atom is bonded to three others arranged at the corners of a regular triangle. Each graphite layer is considered a two-dimensional giant molecule. The layers are held together by weak forces of attraction, enabling the layers to easily slide past each other.

Properties:

- Graphite is soft and slippery (*due to weak forces of attraction between the layers*).
- Graphite conducts electricity (*due to the free delocalized electrons that can move and carry electric current*).
- Graphite has high melting point (*due to the strong covalent bond between carbon atoms within a single layer*).

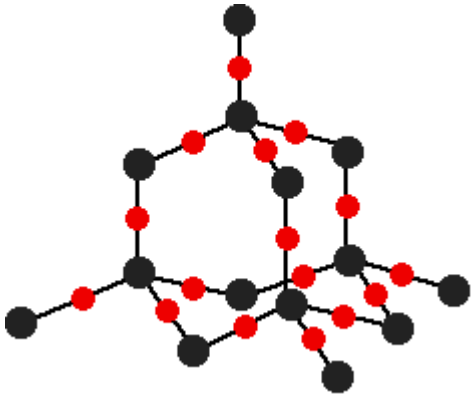

Strong bonds exist between carbon atoms in a layer
Weak forces hold the layers of carbon atoms together
atoms in one layer:

Summary comparing diamond and graphite

Diamond	Graphite
	
Structure Diamond is a tetrahedral structure of linked hexagons, where every carbon atom is strongly and covalently bonded to four other carbon atoms, thus there are no free electrons.	**9. Structure** It is a layered structure of linked hexagons, where every carbon atom is strongly and covalently bonded to three other carbon atoms, leaving one electron unbounded, this forms a cloud of electrons or delocalized electrons within each of the layers.
Properties 1. Diamond has high melting point due to the strong covalent bond between carbon atoms. 2. Diamond does not conduct electricity because it does not have free electron or ions that can move and carry electric current. 3. Diamond is hard due to the strong covalent bonds between carbon atoms.	**Properties** 1. Graphite is soft and slippery due to weak forces of attraction between the layers (VDW). 2. Graphite conducts electricity due to the free delocalized electrons that can move and carry electric current 3. Graphite is hard due to the strong covalent bonds between carbon atoms. 4. Graphite has high melting point due to the strong covalent bond between carbon atoms.
Uses 1. Jewellery (lustre) 2. Glass cutters (hard) 3. Drill bits (hard) 4. Diamond saw (hard) 5. Rock borers (hard) **6.** Abrasive materials (hard)	**Uses** 1. Furnace lining (high melting point) 2. Electrodes (electrical conductor) 3. Lubricants (slippery) 4. Pencils (slippery/ soft)

Giant molecular structure

Silica or Silicon dioxide or Silicon (IV) oxide (for eg. sand and quartz)



Structure

Silicon dioxide has the same structure as diamond

It is a tetrahedral structure where every silicon atom is strongly and covalently bonded to four oxygen atoms and every oxygen atom is strongly and covalently bonded to two silicon atoms leaving no free electrons.

Properties (very similar to diamond due to similar structure)

- Silica does not conduct electricity (*it does not have free electron or ions that can move and carry electric current*).
- High melting point (lower than diamond) because the strength of the covalent bond between silicon and oxygen in silica is weaker than that between the two carbon atoms in diamond.
- Very hard (less hard than diamond) because the strength of the covalent bond between silicon and oxygen in silica is weaker than that between the two carbon atoms in diamond.

Uses In microchips for computers

19.[G] Describe the structure of a network solid and give 2 examples on such solids.

22. [T] Which of the following is **not** a common property for both diamonds and silica?

A. They both form giant covalent structures.
B. They are hard substances.
C. They both have relatively high melting point.
D. Each of their atoms is surrounded by four covalent bonds.
E. They both have a cage-like structure.

CQD2. Explain the following observation: Sulfur melts at a much lower temperature than graphite.

CQD3. Explain why phosphorous is a brittle element while diamond is hard.

Ex5. What particles make up the crystal lattice of a network solid?

Ex6. Which of the following substances are network solids?

a) Carbon in diamond
b) Helium
c) Hydrogen
d) Carbon in graphite
e) Silicon dioxide
f) Quartz

CQD9. a. Germanium has a similar structure to that of diamond. Describe the structure of germanium.

b. Silicon dioxide SiO_2 has a similar structure to that of diamond. Describe the structure of silicon dioxide. You may use a diagram if necessary.

CQD1. Specify the major intermolecular forces that exist in a crystal of each of the following elements: H_2, He, O_2, C, Na, P_4, and S_8.

Element	Major intermolecular forces
H_2	
He	
O_2	
C	
Na	
P_4	
S_8	

Ex7. Specify the major intermolecular forces that exist in a crystal of each of the following elements:

Ne: Cl_2: K: F_2:
Xe: C (diamond): Al:

CQD10. Specify the major forces/bonds that exist in a crystal of each of the following compounds:

CH_4, HCl, CO_2, NH_3, HF:
NaCl:
SiO_2:

Ex8. Specify the major intermolecular forces that exist in a crystal of each of the following compounds:

a) CCl_4, CO_2, NH_3, PH_3, HCl, NaCl, SiO_2, HF, CS_2, CH_4, HBr, HI, C_2H_5OH, NaOH, CaI_2

CCl_4, CO_2, NH_3, PH_3, HCl, HF, CS_2, CH_4, HBr, HI, C_2H_5OH:
NaCl, NaOH, CaI_2:
SiO_2:

Ex9. Which of the following is/are correct properties of a substance that has Van der Waal's forces between its molecules?

a) It is hard
b) It has a low melting point
c) It is a good conductor of electricity
d) It has a low heat of vaporization
e) It has a high boiling point

Ex10.

a) What particles make up the crystal lattice of a molecular substance like hydrogen chloride?
b) Iodine is a molecular solid. A crystal of iodine is made of _____

CQD12. Describe the structure of solid CO_2, I_2, KCl, diamond, H_2O, Li, SiO_2 and Xe. Are they made of ions, giant structures or separate molecules?

Substance	**Type of Solid**	**Particles making up solid**
CO_2		
I_2		
KCl		
diamond		
H_2O		
Li		
Xe		
SiO_2		

CQD13. The following table is a list of the electron distribution of atoms of certain elements.

element	**electron distribution**
A	2, 5
B	2, 8, 4
C	2, 8, 8, 2
D	2, 8, 18, 8
E	2, 8, 18, 8, 1
F	2, 8, 18, 18, 7

a) From the above list, choose an element that:

i) …… is a noble gas.
ii) …… is a soft metal with a low density.
iii) …… can form a covalent compound with element **A**.
iv) …… has a giant molecular structure similar to diamond. **B**
v) …… can forma negative ion of the type X^{3-}. **A**

b) Elements **C** and **F** form an ionic compound.
Draw a diagram that represents the formula of this compound, the arrangement of the valency electrons around the negative ion and the charge on the ions.
Use ◦ to represent an electron from atom **C.**
Use x to represent an electron from atom **F**.

c) Draw a diagram representing the arrangement of valency electrons in the molecule of nitrogen tirchloride.
Use ◦ to represent an electron from chlorine atom.
Use x to represent an electron from nitrogen atom.

CQD23. [T]
The table shows examples on macromolecules, covalent compounds, metals, and ionic compounds with some of their properties.

Substance	Melting point / °C	Boiling point / °C	electrical conductivity when solid	electrical conductivity when liquid	Solubility in water
A	801	1413	non-conductor	good conductor	Soluble
B	825	1339	non-conductor	good conductor	Insoluble
C	-95	69	non-conductor	non-conductor	Insoluble
D	-78.4	-56.6	non-conductor	non-conductor	Soluble
E	1083	2562	good conductor	good conductor	Insoluble
F	1710	2230	non-conductor	non-conductor	Insoluble

a. Which substance is a metal?
b. i. Which substances are ionic compounds? ii. Explain your answer.
c. Which substance conducts electricity in aqueous solution?
d. Which substance is a gas at room temperature?
e. Which substance is a liquid at room temperature?
f. i. Which of the given substances is a network solid?
ii. A student suggests that the network solid in the above table is graphite.
Do you agree with him? Explain your reasoning.

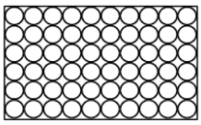
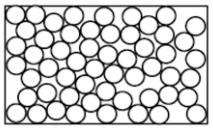
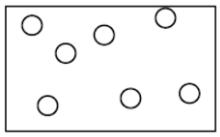
CQD24. [T] Potassium nitride melts at 800^{o}C. Nitrogen trifluoride melts at -200^{o}C.
Explain in terms of attractive forces why potassium nitride has a much higher melting point than nitrogen trifluoride.
In your answer refer to the types of attractive forces between particles and their relative strengths.

Topic 1-Experimental Techniques

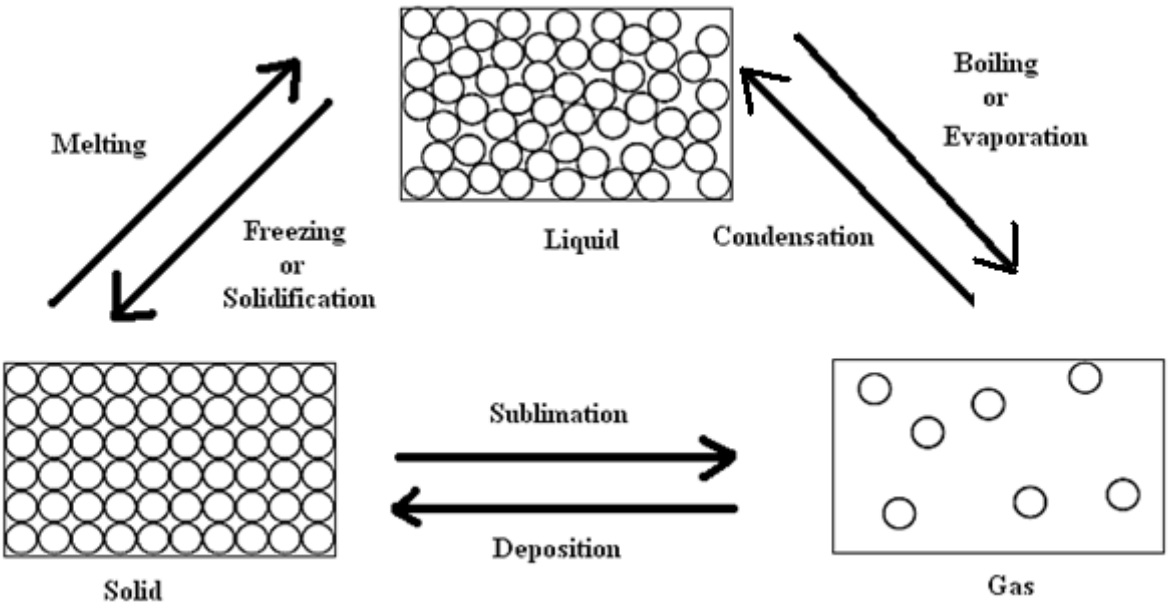
The Particulate Nature of Matter – States of Matter

- There are three states of matter: solid, liquid and gas
- All matter is made up of tiny particles that are too small to be seen by the naked eye.
- The particles are arranged and move in a different way in each state of matter.
- We can change the state of matter by changing the amount of energy its particles have.

Comparison between the three states of matter

	Solid	Liquid	Gas
Particles in substance			
Space between particles	Close together, touching one another	Close together, touching one another	Far apart
Arrangment of particles	Regular, repeating pattern	Irrregular	Irregular
Movement of particles	Vibrate around their fixed position but do not move apart	Move around and slide past each other	Move freely and collide with each other
Forces between particles	Stronger than in liquid	Not as strong as in solid	Non-existant
Shape	Fixed shape	Take the shape of their container	Take the shape of their container
Compression	Cannot be compressed	Cannot be compressed	Can be compressed

Changes of state



Process	Description	Energy invovled
Melting	The change of state from solid to liquid at a definite temperature	Endothermic
Boiling or evaporation	The change of state from liquid to gas at a definite temperature	Endothermic
Freezing or solidification	The change of state from liquid to solid at a definite temperature	Exothermic
Condensation	The change of state from gas to liquid at a definite temperature	Exothermic
Sublimation	The change of state from solid to the gas directly without passing by the liquid state. Example: Solid carbon dioxide, iodine and ammonium chloride	Exothermic

Kinetic theory of matter

- Matter is made up of tiny invisible particles that are in constant motion.
- The lighter the particle the faster it moves.
- As temperature increases, the particles move faster.

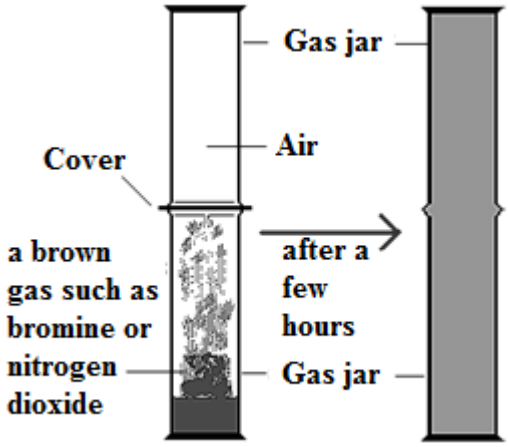
Brownian motion: Particles in fluids (both liquids and gases) move randomly. This is called Brownian motion. They do so because they are *bombarded by the other moving particles in the fluid*. Larger particles can be moved by light, fast-moving molecules.

Brownian motion is named after **Robert Brown**, who first observed this in 1827. He used a microscope to look at pollen grains moving randomly in water. At this point, he could not explain why this occurred. But in 1905, physicist **Albert Einstein** explained that the pollen grains were being moved by individual water molecules. This confirmed that atoms and molecules did exist. Brownian motion is actually an evidence for particle theory.

Diffusion: is the conitinous movement of particles from one place to another to fill the space available for them. Diffusion occurs in both liquids and gases.

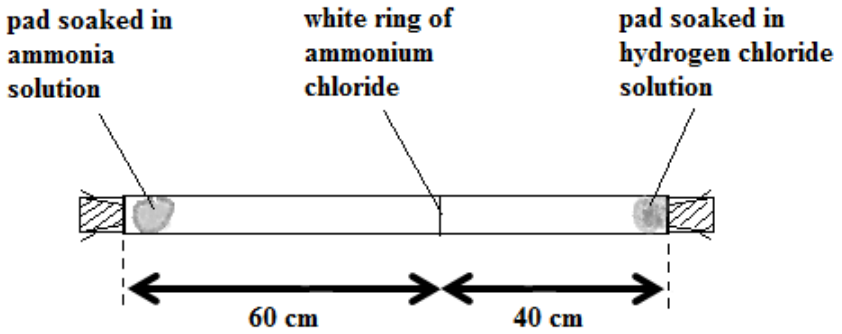
a) Diffusion in gases

Example 1: Diffusion of Bromine (or nitrogen dioxide)



- Two gas jars are placed on top of each other separated by a glass cover.
- The top jar contains air, the bottom jar contains a gas denser than air, for example bromine vapour or nitrogen dioxide gas (used because they are colored).
- If the glass cover is removed, the colorless air moves down and the colored brown gas moves up.
- This is a clear evidence of diffusion due to the random continuous movement of all the gas particles. Initially, the net movement of one type of particle is from a region of higher concentration to a region of lower concentration, colorless gas is moving down and brown gas moving up (down a diffusion gradient). When fully mixed, no further color change distribution is observed **BUT** the random particle movement continues.

Example 2: Formation of ammonium chloride



Ammonia and hydrogen chloride gases react to form a white solid, ammonium chloride:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

- A piece of cotton wool soaked in concentrated ammonia solution (a source of ammonia gas) is placed at the end of a long glass tube.
- A piece of cotton wool soaked in concentrated hydrochloric acid (a source of hydrogen chloride gas) is placed at the other end of the tube.
- The tube is closed with rubber stoppers to stop the gases escaping.
- After few minutes, a white ring of ammonium chloride is formed closer to the hydrochloric acid end.

Although the gas particles move at a speed of several hundred meters per second at room temperature, yet it takes a few minutes for the white ring to form. This is because the gas particles are in random motion, they collide with one another and with the walls of the tube.

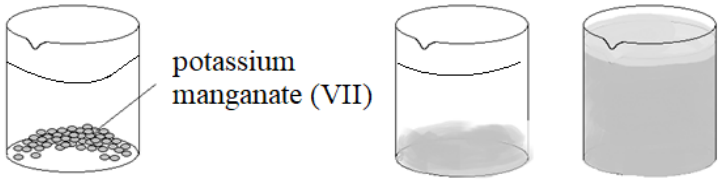
Many collisions are needed for the white ring to become visible.
The white ring is formed closer to the hydrochloric acid end because ammonia particles (17 g/mole) are lighter than hydrogen chloride particles (36.5 g/mole) and so they move faster.

When the experiment was repeated using concenrated hydrobromic acid instead of hydrochloric acid, the white ring formed even closer to the hydrobromic acid side. This is because hydrogen bromide particles (81 g/mole) are much heavier than hydrogen chloride particles and travel much slower.

b) Diffusion in liquids

Diffusion in liquids is much slower than dissusion in gases because liquid particles move much slower than gaseous particles.

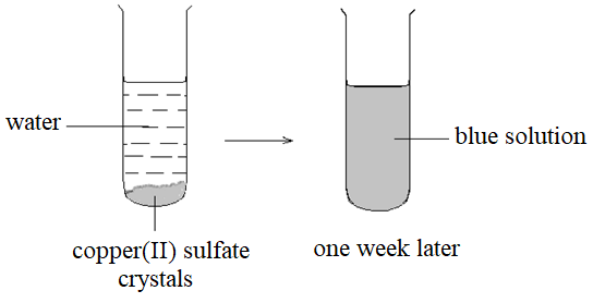
Example 1: Potassium permenganate crystals, $KMnO_4$, in a beaker of water



- Molecules of water collide with the particles of solid $KMnO_4$.
- Potassium permanganate crystals first dissolve then the particles are spread out by the movement of water molecules from region of high concentration to region of low concentration , the color slowly diffuses throughout the liquid.
- After a few days, the particles have mixed completely with water.

Diffusion occurs more rapidly if the liquid is hot. This is because the particles in a hot liquid have more kinetic energy than a cold liquid. These particles move more rapidly, thus spreading out more quickly throughout the liquid.

Example 2: When few blue crystals of copper (II) sulphate is put in water, after one week the water turns into a blue solution



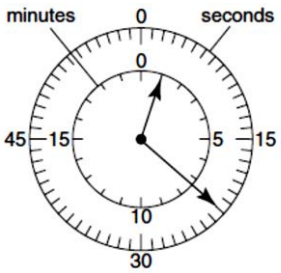
Measurements

During the study of chemistry many experiments are carried out and many measurements need to be taken. It is very important to use the correct appartus for the required task and to be able to read the apparatus correctly.

1. **Time**

   Time is measured using a stopclock or a stopwatch. Expecially important when measuring reaction rates.

   Digital stopwatches read up to two decimal places.

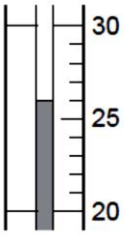


Recorded time: 1 minute and 22 seconds

Reading should be to the smallest division only

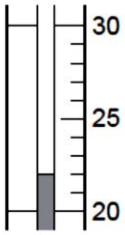
2. **Temperature**

Temperature is generally measured using a thermometer that has intervals of 1°C.



initial temperature

26° C



final temperature

22° C

3. **Mass**

Mass is measured using a top-pan balance.

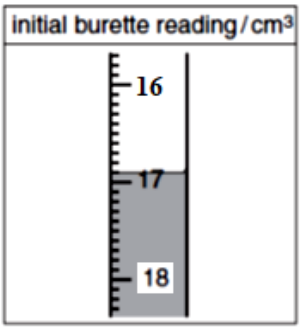
4. **Volume**

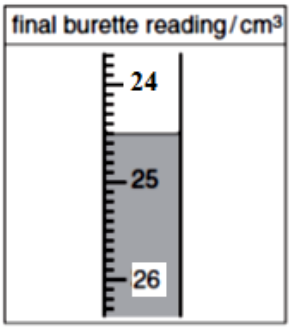
i. Volume of liquids

The volume of liquid is measured with burettes, volumetric pipettes or measuring cylinders depending on the accuracy and the quantities needed.

a) Burette

It is used to measure variable volumes of liquids up to 50 cm^3.





16.9 cm^3 **24.5 cm^3**

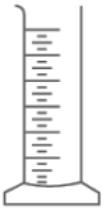
A burette can be used to measure the volume of a gas.

b) Pipette

It is used to measure accurate volumes of liquids.
The common capcaities of a pipette are: 5, 10, 25 and 50 cm^3.
A pipette cannot be used to measure the volume of a gas as it is not graduated, it has one calibration mark.

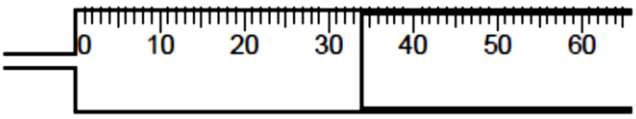
c) Measuring cylinder

It is used to measure approximate volumes of liquids.
A measuring cylinder can be used to collect and measure the volume of a gas.



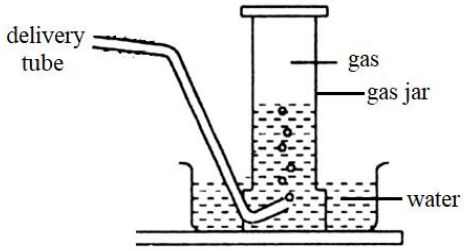
ii. Volume of gases

The volume of a gas is measured most accurately using a gas syringe.



34 cm^3

Note: A gas can also be ***collected*** using a gas jar.
A gas jar cannot ***measure*** the volume of a gas as it is not graduated like a measuring cylinder.



5. pH

It is meaured using a pH meter or Universal indicator.

Reliability

Data collected during an experimant needs to be repeatable and reproducible.
Repeatable: the same person does the experiment several times under the same conditions and gets similar results every time.
Reproducible: different people do the experiment under under the same conditions and get similar results every time.

To make sure that an experiment is repeatable and reproducible the variables need to studied while planning the experiment. There are three main variables to consider:

- controlled variable: the variable kept constant during an experiment
- independent variable: the variable that is changed during an experiment (always plotted on the x-axis)
- dependant variable: the variable measured during an experiment (always plotted on the y-axis)

Sources of error

Every measurement has some degree of error or uncertainty.

- Random error: unpredictable changes in results like human errors
- Systematic error: consistant error (most probably due to problems with the experimental design or defective apparatus)
- Anomalos: something unusual or unexpected (like a measurement too high or too low as compared to the trend or norm)
- Zero error: a type of systematic error, for example an instrment does not read zero when not in use (for eg., balance not reading zero even with nothing on it)

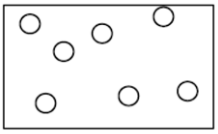
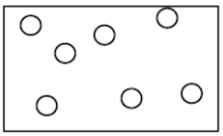
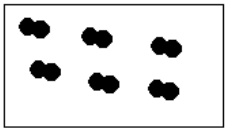
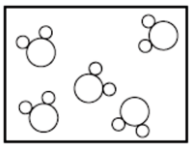
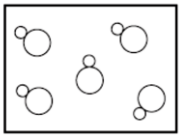
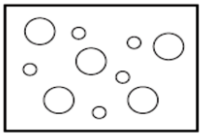
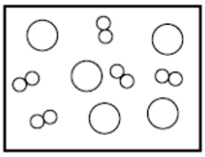
Graph plotting:

- independent variable must be plotted on the x-axis
- dependant variable must be plotted on the y-axis
- make scales as large as possible (they need not start at zero)
- label axis with variable name AND its unit
- Give your graph a title
- Plot the points using a cross (not dot) using a sharp pencil
- Draw the line of best fit. It may be a straight line or a smooth curve. It needs not pass through all points, but try to draw it so that the points are evenly scattered on either side of line.

Safety

When planning an experiment it is essential to assess risks and identify sources of danger to eliminate them.

Elements, Compounds and mixtures

Atom The smallest particle of matter *(The smallest particle of an element which shows the properties of that element. Atoms have equal number of electrons and protons and are electrically neutral)*	Example: Helium atoms, He 
Element • A pure substance which is made from only one type of atoms. • It cannot be broken down into simpler substances by chemical processes *An element contains atoms with the same atomic number*	Examples: Neon, Ne  Oxygen gas, O_2 
Compound *A substance made up of two or more elements which are chemically combined together.* • A compound is a completely new material that will have totally different properties from the elements that made it. • A compound cannot be separated by physical methods.	Examples Water, H_2O  Hydrogen chloride, HCl 
Mixture *A collection of at least two or more elements or compounds which are not chemically bonded together.* • A mixture can be separated by physical methods.	Examples Mixture of helium,He, and argon atoms, Ar  Mixture of hydrogen molecules, H_2, and helium atoms 

CQD 1. Define an element, a compound, and a mixture.

CQD 2. Classify the following as elements, compounds, or mixtures.

Chemical	Element	Compound	Mixture
Blood			
Oxygen			
Ammonia			
Orange juice			

Words to remember:

Solvent: substance that dissolves a solute
Solute: substance that dissolves in a solvent
Solution: a homogenous mixture, a mixture of one or more solutes dissolved in a solvent
Saturated solution: a solution containing the maximum concentration of a solute dissolved in the solvent at a specified temperature
Residue: a substance that remains after evaporation, distillation, filtration or any similar process
Filtrate: a liquid or solution that has passed through a filter

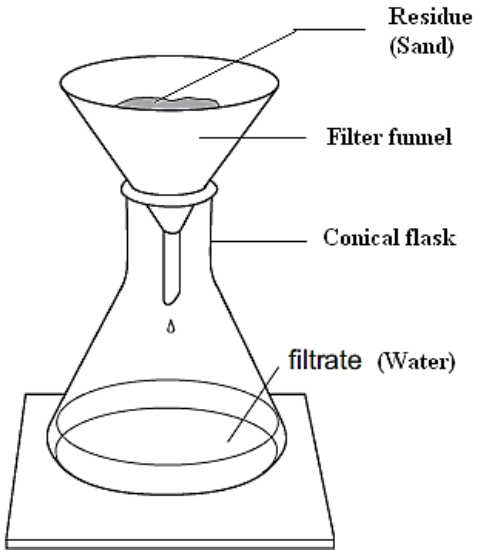
Separation Techniques
The separation technique chosen depends on the type of substances that need separating.
The key to do this is to:

- look for differences in the properties of the components of the mixture such as their state, solubility, melting/ boiling point and so on.
- Then pick the method that best takes advantage of this difference.

1. **Filtration**
   This method is used to separate an **insoluble** solid from a mixture of the solid and liquid.

Example: Separation of a mixture of sand and water

- The mixture of the solid and liquid is poured into a filter paper in a funnel (as shown below).
- The **filtrate** flows through the filter paper and into the flask.
- The solid **residue** remains in the filter paper.
- The residue is washed with a little more solvent
- The residue is dried by spreading it out on the filter paper and either allowing the liquid to evaporate on its own or placing paper in an oven.



2. **Suitable Solvent**

This method can be used to separate to solid mixture if a solvent can be found that dissolves only one of the components.

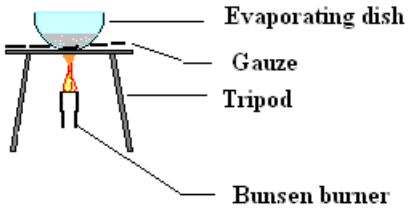
Example: Separation of salt and sugar

- Find suitable solvent: sugar dissolves in alcohol, salt does not
- Dissolve the mixture in minimum volume of alcohol
- Filter the salt out, wash it with little alcohol and leave it to dry.
- Gently heat the filtrate over a water bath heated by an electric burner till the liquid evaporates.

3. **Crystallization**

This method is used to separate a solid which dissolved in a **solution**.

- Heat the solution in an evaporating dish directly over a flame till most of the liquid evaporates.



- Transfer the evaporating dish over a steam bath. Heat gently till the crystallization point (until saturation)

Crystallization point or saturation point is checked by dipping a cold glass rod in the solution from time to time and checking for crystals formation.

- Leave to cool.
- Filter off the crystals.
- Dry the crystals between two filter papers without pressing / or in open air.

CQD 3. Briefly describe how to separate a mixture of a salt and sand into its constituents.

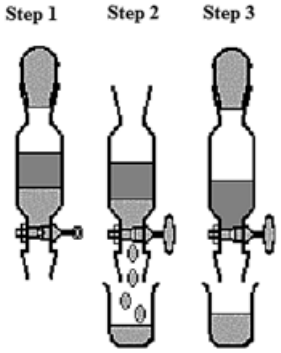
CQD 5. Give the name of an apparatus or the name of a technique you can use to separate the following mixtures into their constituents:

a. A mixture of sulfur and steel.
b. A mixture of cooking oil and water.

CQD 6 [G] Name the most suitable technique used to obtain large potassium nitrate crystals from its aqueous solution. Then, describe how to conduct this technique.

4. **Separating funnel**

A separating funnel is used to separate two immiscible liquids (liquids that do not mix in each other.) A separating funnel is a glass funnel with a tap at the bottom.



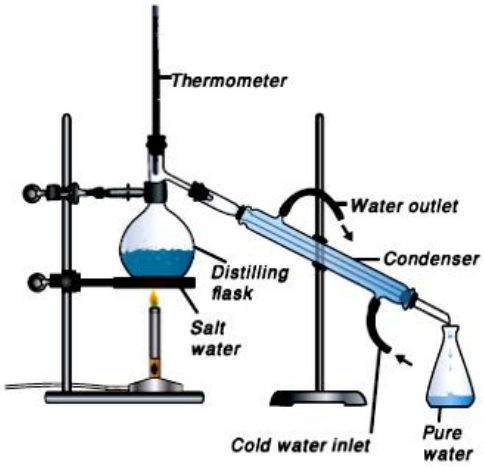
- The liquid mixture is placed inside a separating funnel and left to separate.
- The liquid with the lower density floats on top.
- When the tap is opened, the liquid with the higher density starts to flow down into the collecting container.
- The tap is closed just before the liquid with the lower density starts to flow through.
- The liquid with the lower density remaining in the separating funnel can then be drained into a different container to separate the two liquids.

5. **Simple distillation**

This method is used to separate a pure liquid from a solution.

Example: Separation of pure water from an aqueous solution of sodium chloride.

- The solution is heated until the liquid with the lowest boiling point (most volatile) boils.
- The vapours travel up, enter the condenser where they condense on the cold surface of the condenser and the liquid is trickles into collection flask.
- The condenser has cold water running through a jacket around the outside to keep the temperature below the boiling point of the vapor. Note that water enters the condenser from the inlet closest to collection flask.
- The collected liquid is called the **distillate**, the liquid left is called the **residue**.



Distillation involves two processes: evaporation followed by condensation

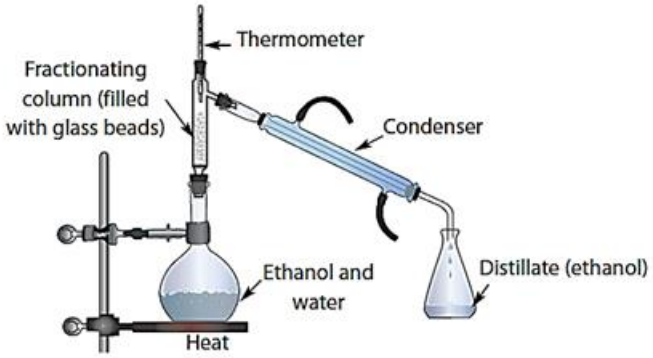
CQD 7. [G] Which of the following mixture(s) is/are best separated by simple distillation?

a. sodium bromide aqueous solution
b. sand, salt, and water mixture
c. water-alcohol solution
d. sugar and water solution
e. any mixture of miscible liquids having close boiling points

6. **Fractional distillation**
   This method is used to separate two or more miscible liquids with close boiling points.

Examples: Separation of ethanol during fermentation
Fermentation is the production of ethanol. It occurs when yeast breaks down sugars in the absence of oxygen. Fermentation is used to make alcoholic drinks, such as beer and wine. Carbon dioxide gas is also produced during fermentation:

glucose (sugar) → ethanol (alcohol) + carbon dioxide



- The solution is heated until the component with the lowest boiling point (most volatile) boils.
- The vapors travel up, enter the condenser where they condense on the cold surface of the condenser and the liquid is trickles into collection flask.

- Once all of the most volatile component is collected, the next lowest boiling point component starts to boil and its vapors enter the fractionating column, condense back to liquid in the condenser and gets collected.
- This continues till all components are separated

Note:
It is not always possible to separate the components 100%.
In this case, the collected distillate will have a higher percentage composition of the more volatile component.

CQD 8. [G] Which of the following mixtures is best separated by fractional distillation?

a. ethanol-water solution
b. oil-water mixture
c. salt water solution
d. sugar alcohol solution
e. any mixture of miscible liquids having close boiling points

7. Chromatography

There are many different kinds of chromatography, but they all depend on the same principles. The mixture is dissolved in a fluid (called the mobile phase or solvent) and allowed to pass through a second material (called the stationary phase).

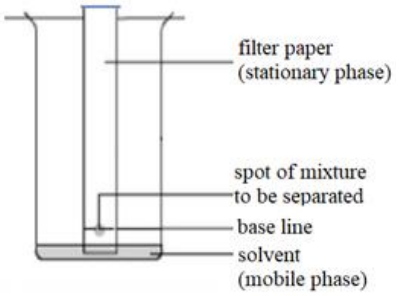
Paper chromatography

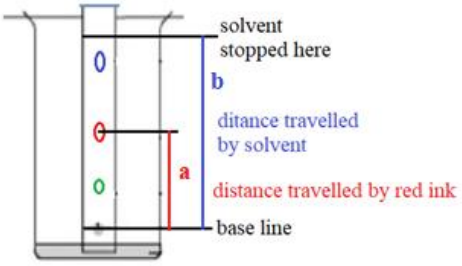
This method can be used to separate mixtures of

- soluble colored substances using a suitable solvent, eg. components of chlorophyll, components of ink, components of food colorings.
- Soluble colorless substances using a suitable solvent and a locating agent, eg. amino acids from a mixture of amino acids

As such, paper chromatography is extensively used in the food industry and forensic science.

Example: Separation of black ink into its components.





$$R_f \text{ for red dye} = \frac{a}{b}$$

- Black ink is placed as a small concentrated dot on a filter paper (stationary phase) on a line that is drawn in pencil, called base line. This line is drawn in pencil, not pen, because the colored components in the pen will interfere with the experiment.
- The paper is then placed in a suitable solvent like water, alcohol or a mixture of both (mobile phase). The level of the solvent must be below the base line as dyes might wash out directly in the solvent. If alcohol is used the beaker must be covered with a watch glass to prevent the loss of alcohol.
- Solvent starts rising up the paper by capillary action. As it rises up, the components (dyes) within the ink spots are separated.
- When the solvent is about 1 cm from the top of the paper, the paper is removed.
- Draw another line in pencil to show where the solvent reached. This line is called the solvent front.
- Different components have different solubility in the solvent and different molecular adhesion to the paper. As such, different components travel different distances up the paper allowing the separation of the different components.
- The component that is most soluble in the solvent and adheres least to the paper will travel the longest distance up the paper. Conversely, the component that is least soluble in the solvent and adheres most to the paper will travel the shortest distance up the paper and so on.
- The resulting paper with the separated components is called a chromatogram.

CQD 4. What process can be used to separate the dyes in food coloring?

Sometimes the components of a mixture are colorless, *e.g amino acids mixture*, nothing will be visible on the chromatogram. In that case, the chromatogram can be sprayed by a suitable locating agent. A locating agent interacts with the components coloring them so they become visible.

What is the R_f value?

R_f value or retention factor is the distance travelled by a component divided by the distance travelled by the solvent, measured from the baseline.

$$R_f = \frac{distance\ travelled\ by\ substance}{distance\ travelled\ by\ solvent}$$

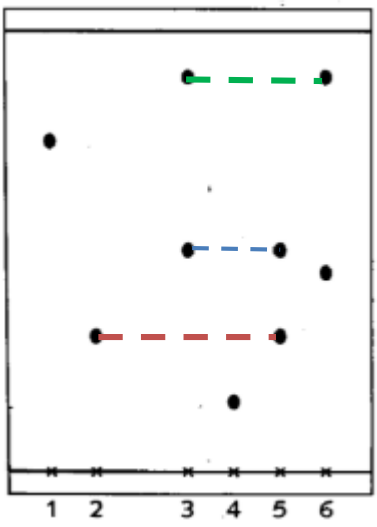
R_f is a physical property, different substances have unique R_f values therefore R_f values van be used to identify substances.

In addition to separating and identifying substances in a solution, chromatography can also be used to check the purity of a given substance *(see interpreting chromatograms)*

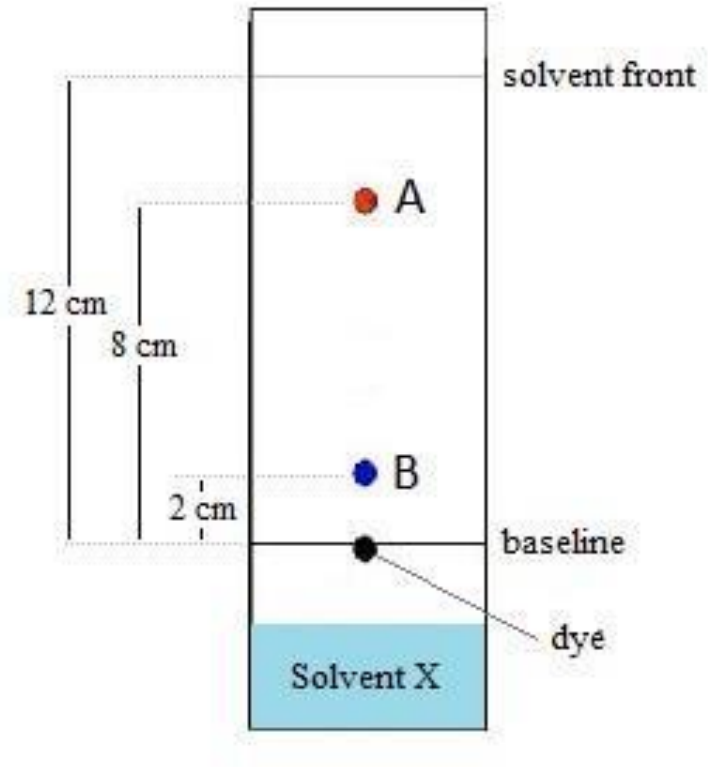
Interpreting chromatograms:

Consider the following chromatogram of six substances.

- Substances 1, 2 and 4 are soluble pure substances because the initial spot moved up the paper with the solvent and did not split.
- Dyes 3, 5 and 6 are mixtures of soluble components because the initial spot moved up with the solvent and split into more than one spot. Remember that the greater the solubility of a component the greater is its R_f value the larger is the separation from the base line.
- Dyes **2 and 5**, **3 and 5**, **3 and 6** have a common component (same R_f value).
- If a substance does not move up and remains on the base line, it is considered to be insoluble in the solvent.



CQD 9. [G] The following shows the chromatogram of a dye that was separated into two components A and B using paper chromatography in a solvent X.



a. Which of the two components of the dye is more soluble in solvent X and has the least affinity to the paper?
b. Find the R_f value for component A.

CQD 10. [G] A mixture of amino acids is a colorless solution and can be separated using chromatography. Describe how the chromatography procedure should be adapted to reveal the amino acids spots on the chromatogram.

The purity of solids and liquids

The best way to check the purity of solids is to determine its melting point.
The best way to check the purity of liquids is to determine its boiling point.

Pure substances have sharp, fixed melting and boiling points.
An impure solid will have a melting point **lower** than the pure solid.
Impurities decrease melting points.
An impure liquid will have a **higher** boiling point than the pure liquid.
Impurities increase boiling points.

Examples:
Pure water boils at 100℃ while salt water (sodium chloride in water) boils at 102℃
Pure ice melts at 0℃ while ice that has salt in it melts at about - 4℃.

Chromatography can be used to determine the purity for colored dyes as we discussed earlier.

CQD 11. You are given a sample of colorless liquid in a test tube and are told that this is water. How can you assess the purity of the given sample?

Methods of Purification

1. **Purification of impure solids**

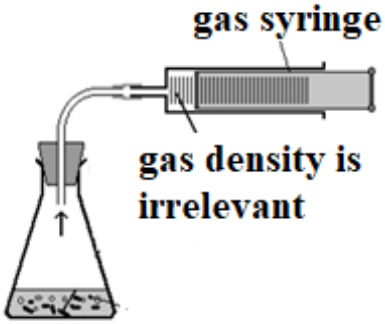
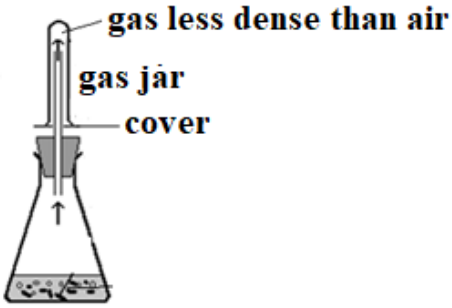
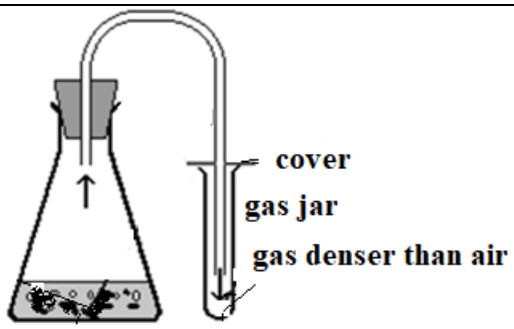
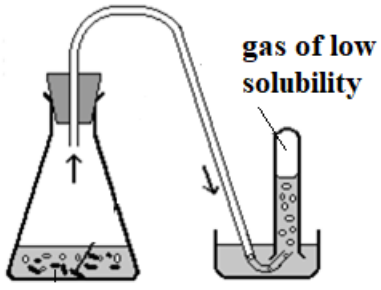
Example: To separate salt from rock salt.

- Dissolve the solid in a solvent in which it is soluble.
- Filter the mixture to remove any insoluble substances (impurities) as a residue.
- Evaporate the solution to reduce its volume then leave the solution to crystallize (refer to method 1)
- Filter off the crystals leave the crystals to dry. The result is a pure solid.

2. **Purification of impure liquids**
    - Simple distillation
    - Fractional distillation

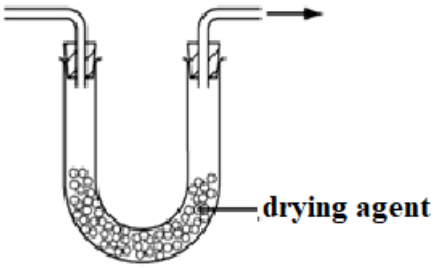
Working with gases

(1) Collection of gases

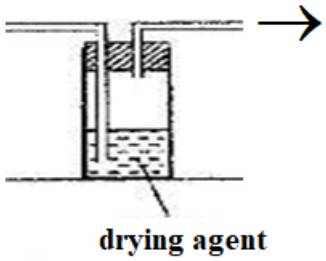
Method	Diagram	Examples
Any gas can be collected by gas syringe		Any gas
Gases **less** dense than air are collected by upward delivery.		Hydrogen, H_2 Ammonia, NH_3
Gases **denser** than air are collected by downward delivery		Chlorine, Cl_2 Sulphur dioxide, SO_2 Hydrogen sulphide, H_2S Nitrogen dioxide, NO_2 Carbon dioxide, CO_2 Hydrogen chloride, HCl
Gases collected over water (gas must be insoluble in water)		Oxygen, O_2 Hydrogen, H_2 Nitrogen, N_2

(2) Drying of gases
Pass the gas through a drying agent. Drying agents can be solids or liquids.

Example of solid drying agent: Anhydrous calcium chloride solid



Example of liquid drying agent: Concentrated sulfuric acid



Notes:

- toxic (poisonous) gases like chlorine / pollutant gases like nitrogen dioxide should be prepared in a fume cupboard as a precautionary measure.
- Conc. Sulfuric acid is never used to dry ammonia gas, because ammonia gas is basic and sulfuric acid is acidic so they will react together.
- The most suitable drying agent for ammonia gas is calcium oxide solid.

CQD 12.
a. Calculate the molecular masses of H_2S and NH_3. Compare their densities under identical conditions. **Report your answer as a simple fraction.**
b. Which of the above two gases is collected by upward delivery, and which is collected by downward delivery?

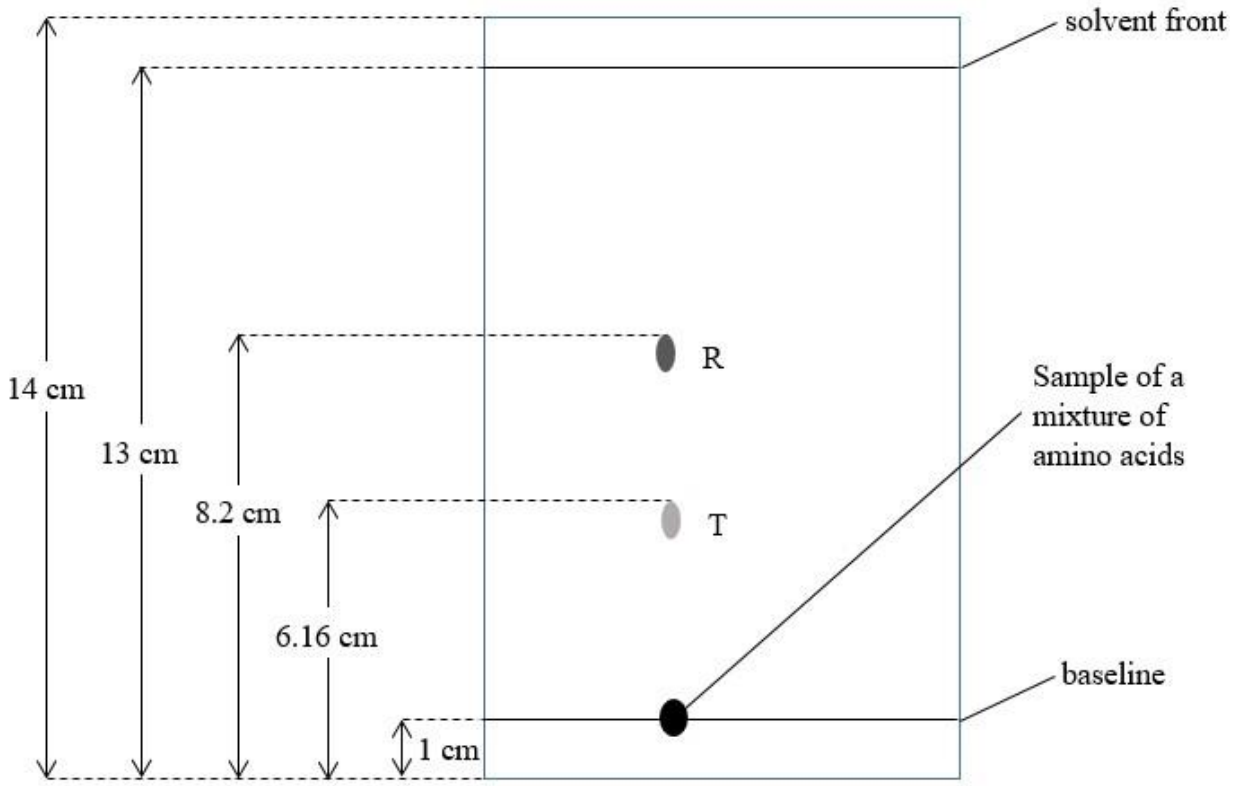
CQD 13. Name two common drying agents that are used to dry gases.

CQD 14.
a. Name three gases that are highly soluble in water.
b. Draw a labeled diagram of a setup you can use to make an aqueous solution of a highly soluble gas in water.

CQD 15. [T] The table below provides the R_f values of four amino acids.

Amino acid	R_f value
lysine	0.14
glycine	0.26
proline	0.43
valine	0.60

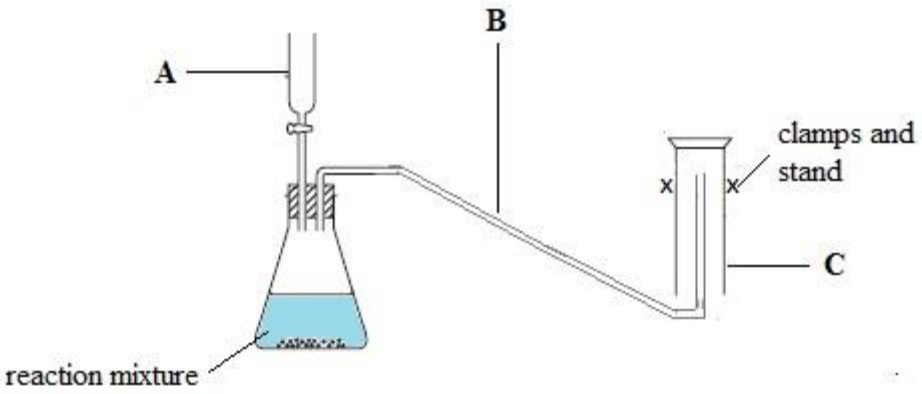
The following diagram shows a chromatogram for a mixture of two amino acids R and T.



What are the identities of the amino acids separated?

	R	T
A	proline	valine
B	lysine	proline
C	valine	glycine
D	valine	proline
E	glycine	proline

CQD 16. [T] The apparatus shown is used to collect gases in the lab.



a. Name the apparatus **A**, **B**, and **C**.

b. A student used this apparatus to collect sulfur dioxide gas.

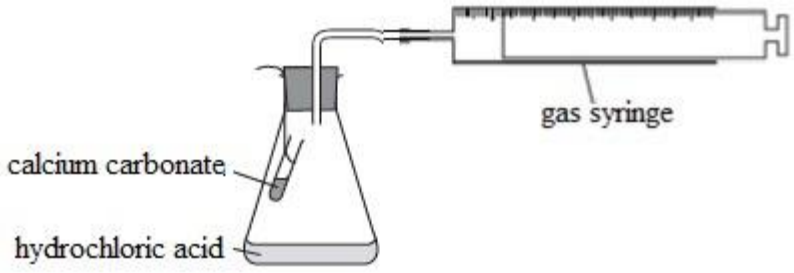
i. Explain why the student made a mistake using this apparatus to collect sulfur dioxide gas.

ii. Name one gas that can be collected using this apparatus.

c. Suggest a modification on the apparatus to collect sulfur dioxide gas. You may include a diagram in your answer.

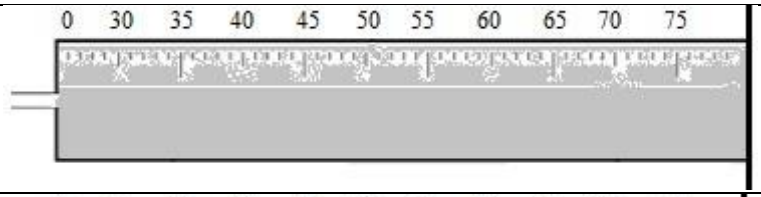
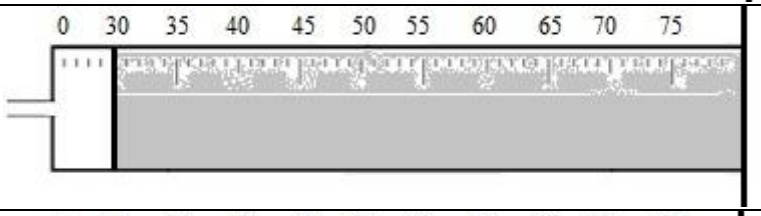
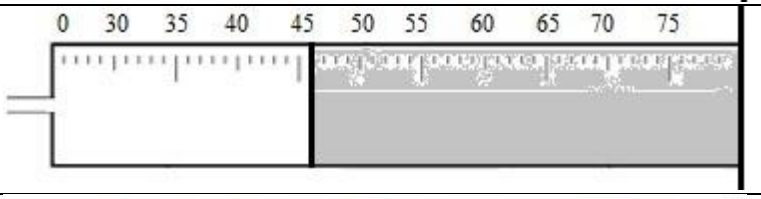
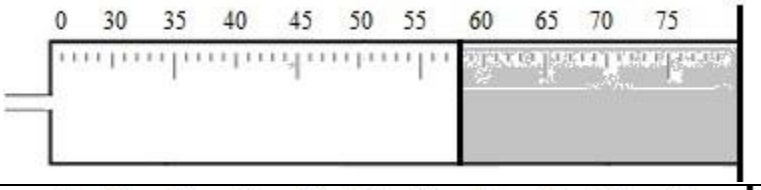
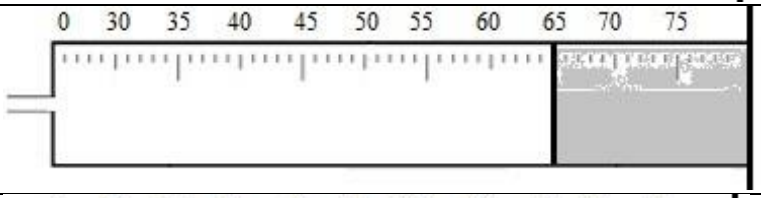
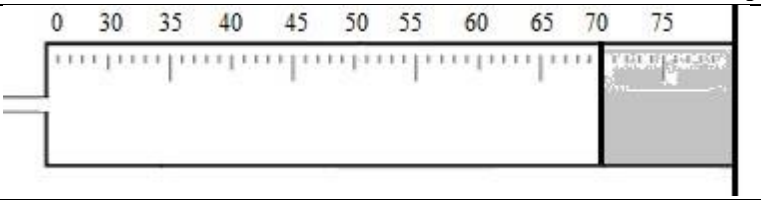
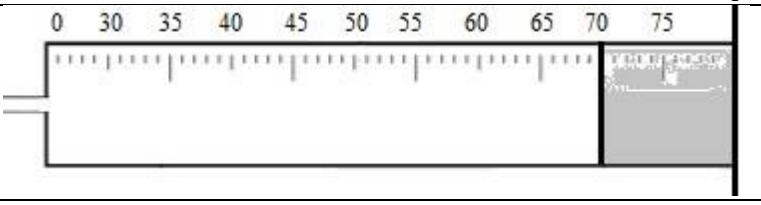
CQD 17. [T] A student investigated the rate of reaction between dilute hydrochloric acid and excess calcium carbonate.

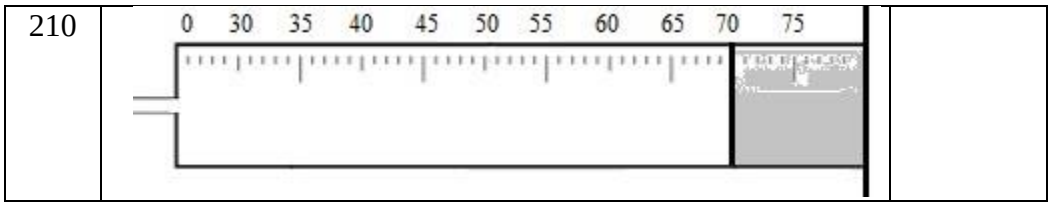
The student used a gas syringe to collect and measure the amount of gas released as shown in the diagram.



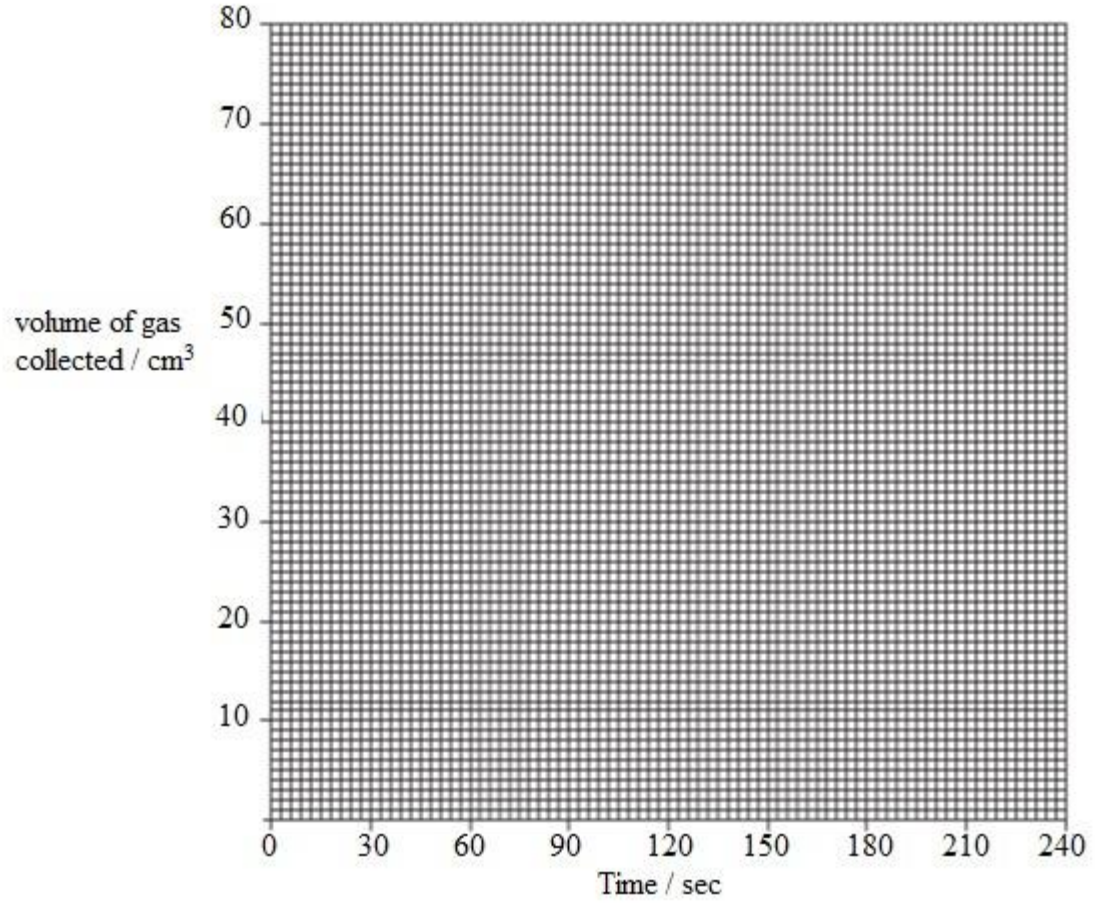
As the calcium carbonate sample was dropped into the hydrochloric acid solution, the volume of carbon dioxide gas was monitored every 30 seconds 25°C.

a. Use the syringe diagram to record the volume of carbon dioxide gas released.

Time / sec	Syringe diagram	Total volume of gas collected / cm^3
0		
30		
60		
90		
120		
150		
180		



b. Plot the results on the grid and draw a smooth line curve.



c. Use the graph and deduce the volume of hydrogen gas released after 80 seconds. Show clearly on the graph how you deduced your answer.

d. If instead of using a gas syringe, another student collected the gas using an inverted 50 cm^3 measuring cylinder filled with water.

i. Suggest an error in this experiment.

ii. Suggest an improvement to fix the error (other than using a gas syringe).

e. The following expression is used to calculate the average rate of the reaction.

average rate of reaction = $\frac{volume\ of\ gas\ collected\ /cm^3}{time/sec}$

i. Calculate the volume of hydrogen gas collected between 30 seconds and 120 seconds.

ii. Calculate the average rate between 30 seconds and 120 seconds. Give your answer to two significant figures and include the unit.

f. Sketch on the same graph the expected curve if the concentration of hydrochloric acid was increased.

g. If the gas to be collected was hydrogen rather than carbon dioxide, would it have been possible to collect it over water using a 250 cm^3 measuring cylinder? Explain.

Topic 3-Water

Water is the most abundant substance on the surface of our planet and is essential for all living organisms

Implication of inadequate supply of water

Unavailability of fresh water leads to:

- waterborne diseases, such as cholera and typhoid
- diarrhea (one of the biggest killers across the world)
- famine, agriculture needs a water supply in order to irrigate crops

Natural Water

Water from natural sources contains many substances, some of which include:

- dissolved oxygen
- metal compounds
- plastics
- sewage
- harmful microbes
- nitrates from fertilisers
- phosphates from fertilisers and detergents

Some of these substances are beneficial to living organisms, like:

- dissolved oxygen for aquatic life
- some metal compounds provide essential minerals for life

While others are potential harmful:

- some metal compounds are toxic
- some plastics harm aquatic life
- sewage contains harmful microbes which cause disease
- nitrates and phosphates lead to deoxygenation of water and damage to aquatic life

Further sources of pollution include:

- acid rain.
- Power stations which generate electricity consume huge amounts of water for cooling purposes. When water is returned to rivers or lakes it will be hotter, decreasing oxygen solubility. Aquatic life suffers as a result of lesser oxygen.

CQD1. Name one natural phenomenon that can cause water pollution.

CQD2. [G] Name the major water pollutants, and classify them as organic and inorganic.

CQD8. Water from natural resources may contain beneficial and harmful substances. Fill the table to list the substances as beneficial and partially harmful.

Substances	Beneficial	Potentially harmful
Dissolved Oxygen	Dissolved Oxygen for aquatic life	
Metal Compounds	Some metal compounds provide essential minerals for life	Some metal compounds are toxic
Sewage Harmful microbes		Sewage contains harmful microbes which cause diseases
Nitrates from fertilizers Phosphates from fertilizers and detergents		Nitrates and phosphates lead to deoxygenation of water & damage to aquatic life
Plastics		Some plastics harm aquatic life

Water Treatment

Steps involved in water treatment are:

Sedimentation, filtration and screening to remove solids
Sterilization to kill microbes
Removal of tastes and odours by using carbon

Before treatment, water has both soluble and insoluble impurities. It is made safe to drink by removing the solid impurities and killing bacteria.

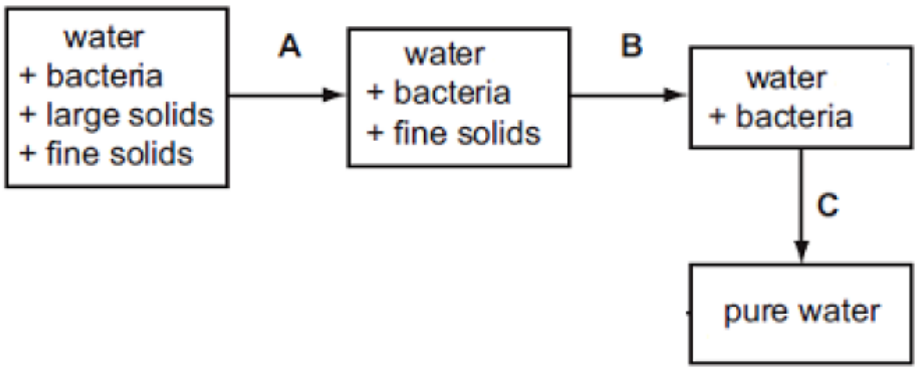
- Solid impurities, like soil, plant parts, organic matter are removed by sedimentation and filtration. Water is stored undisturbed so the solids settle at the bottom (sedimentation). The water is then passed through stone and sand filter beds to remove any solid impurities that escaped sedimentation.
- Chlorine is bubbled through the water to kill bacteria.
- The safe clear water is pumped through underground pipes to houses.

In addition to chlorine, other methods can be used to sterilize water, these include:

- Ultra violet light
- Adding Ozone, the addition of ozone kills the bacteria easily and can also attack some of the organic matter responsible for the bad smell of water. Ozone is a highly poisonous gas so care has to be taken when adding it.

CQD3. Briefly state how water is treated.

CQD4. [G] Polluted water is treated through a process of several purification steps in order to be used for domestic purposes. The following schematic diagram represent the water treatment process.



a. At which stage screening is performed? Explain.
b. At which stage filtration is performed? Explain.
c. Specify between which two steps sedimentation should be performed.
d. What do we call the last stage? Explain.
e. What is used to eliminate odor and taste of water? At which stage is it added?

CQD11. [T] Polluted water undergoes many steps, in a water treatment, before it becomes suitable for drinking. Which of the following steps is used to remove insoluble impurities from polluted water?

A. filtration
B. coagulation C. addition of chlorine
D. aeration
E. none of the above

Removing tastes and odours from water is an important step in water treatment. This is usually done by:

- Aeration, as water is mixed with air any volatile impurities responsible for odours are released
- Filtration through activated carbon, this removes tastes and odours.

CQD 5. What substance can be used to remove tastes and odors in the treatment of water?

Some countries add sodium fluoride, NaF, to the water to reduce the tooth decay. This is not highly recommended because it leads to the appearance of permanent black lines in the teeth.

Test for the presence of water

A hydrated substance is one that is chemically combined with water while an anhydrous substance is one containing no water. The term water of crystallisation refers to the water molecules present in hydrated crystals. Of the anhydrous salts two are very useful to indicate the presence of water: anhydrous copper(II) sulfate and anhydrous cobalt(II) chloride.

- Using anhydrous copper (II) sulfate: Water turns white anhydrous copper (II) sulfate blue.

$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O$
White blue

- Using anhydrous cobalt (II) chloride: Water turns blue anhydrous cobalt (II) chloride pink.

$CoCl_2(s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O$

Blue pink

CQD6. [G] Fill in the blanks to determine the color change upon hydration of given salts.
a. When few drops of water are added to anhydrous copper(II) sulfate, the observed color change is from ____a_____ to _____b_____
b. When few drops of water are added to anhydrous cobalt(II) chloride, the observed color change is from ____c_____ to _____d_____

Test for water purity

By determining the boiling point (100°C) and/or freezing point (0°C).
Any pure material must have a sharp, fixed and definite freezing point and boiling point.
So, if the liquid boils at a fixed 100°C and/or freezes at a fixed 0°C then it is pure water.

Pure water can only be produced by distillation.

- It is not good to drink distilled water. Water is a source of ions for the body, without which health problems may arise.
- It is used in practical chemistry because it contains fewer chemical impurities

CQD 7.
a. How many types of water are there?
b. Describe how to test the purity of water?
c. Why distilled water is used in laboratory experiments rather than tap water?

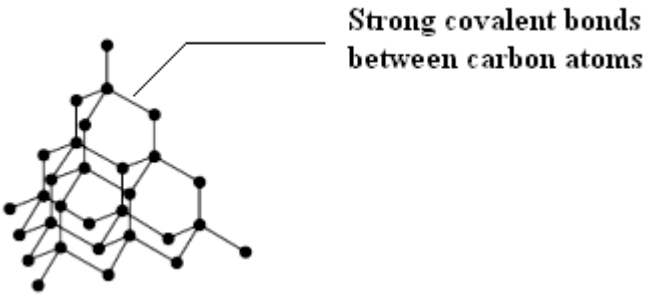
Carbon and its compounds

Carbon

- **Occurrence of carbon**
  Carbon is can be found in a large number of naturally occurring substances such as coal, rocks, fossil fuels, carbonates, proteins, carbohydrates, fats, …. It is also found in very small amounts in air as carbon dioxide (around 0.03 % by volume).

- **Allotropy of carbon**
  Carbon has three known allotropes: diamond, graphite and Buckministerfullerene.

Diamond



Structure

Diamond is a tetrahedral structure of linked hexagons, where every carbon atom is strongly and covalently bonded to four other carbon atoms, thus there are no free electrons in the structure.

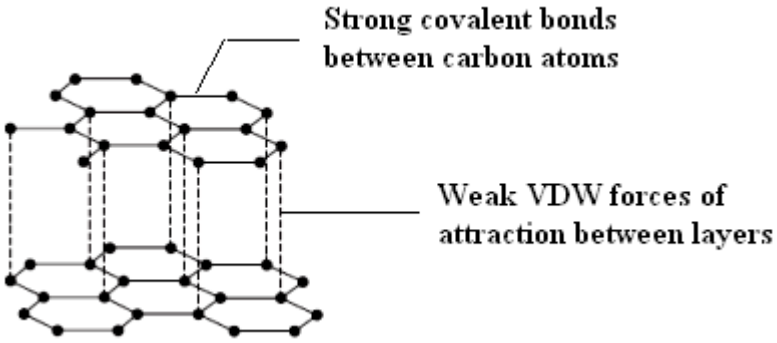
Properties

- Diamond has high melting point due to the strong covalent bond between carbon atoms.
- Diamond does not conduct electricity because it does not have free electron or ions that can move and carry electric current.
- Diamond is hard due to the strong covalent bonds between carbon atoms.

Uses

- Jewellery (due to its lustre)
- Cutting tools (due to its hardness)
- Rock borers (due to its hardness)
- Abrasive materials (due to its hardness)

Graphite



Structure

It is a layered structure of linked hexagons, where every carbon atom is strongly and covalently bonded to three other carbon atoms, leaving one electron free. The free electrons form a cloud of delocalized electrons between the layers.

Properties

- Graphite is soft and slippery due to weak forces of attraction between the layers (VDW).
- Graphite conducts electricity due to the free delocalized electrons that can move and carry electric current within each layer.
- Graphite has high melting point due to the strong covalent bond between carbon atoms.

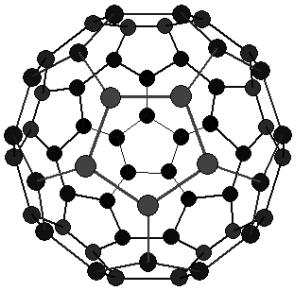
Uses

- Furnace lining (due to its high melting point)
- Electrodes (electrical conductor due to its delocalized electrons)
- Lubricants (slippery, VDW forces between layers are easy to break)
- Pencils (slippery/ soft, VDW forces between layers are easy to break)

Q Compare between the structure of diamond and graphite. Draw a diagram for each.

Q Graphite is soft and conducts electricity, while diamond is hard and does not conduct electricity. Explain.

Buckministerfullerene



Structure

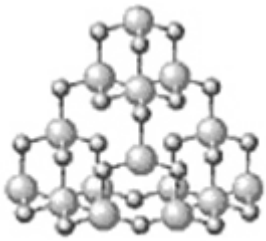
It is molecular with each molecule consisting of 60 atoms. Each carbon atom forms three bonds.

Q: How is fullerene different from graphite and diamond? **It is molecular.**

Silicon

It belongs to the same group as carbon. Consequently, it has similar properties to it. For example, it forms silica whose structure is very similar to diamond.

Silica or Silicon dioxide or Silicon (IV) oxide or Sand



Structure

It is a tetrahedral structure of linked hexagons where every silicon atom is strongly and covalently bonded to four oxygen atoms and every oxygen atom is strongly and covalently bonded to two silicon atoms leaving no free electrons.

Properties

- Silica does not conduct electricity because it does not have free electron or ions that can move and carry electric current.
- High melting point (lower than diamond) because the covalent bond between silicon and oxygen in silica is weaker than that between the two carbon atoms in diamond.
- Very hard (less than diamond) because the covalent bond between silicon and oxygen in silica is weaker than that between the two carbon atoms in diamond.

Uses In microchips for computers

Reactions of carbon

1. **Combustion**
   a) Incomplete combustion of carbon which results in production of carbon monoxide.
      $2C(s) + O_2 \rightarrow 2CO(g)$
   b) Complete combustion of carbon which results in production of carbon dioxide.
      $C(s) + O_2 \rightarrow CO_2(g)$

The process of complete combustion produces more energy per mole than the process of incomplete combustion.

2. **Reducing properties**
   In some of its reactions, carbon acts as a reducing agent.
   $C(s) + O_2(g) \rightarrow CO_2(g)$
   $C(s) + 2CuO(s) \rightarrow 2Cu(s) + CO_2(g)$

Carbon dioxide

Sources of carbon dioxide

- Product of respiration: $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
- Decomposition of animals and plants.
- Complete combustion of fossil fuels: $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
- By the reaction between an acid and a carbonate or hydrogen carbonate.
  $CaCO_3(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$
- Thermal decomposition of metals carbonates

When a metal carbonate is heated, a colorless gas (carbon dioxide) is produced.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\quad\quad$ $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

White soild white solid $\quad\quad$ green solid black solid

Test for carbon dioxide

Test: Pass gas through clear lime water for a short time
Result: Lime water turns milky or cloudy.
Equation: $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

If the gas is passed through lime water for a long time, the turbidity disappears as soluble calcium hydrogen carbonate forms: $CaCO_3(s) + H_2O(l) + CO(g) \rightarrow Ca(HCO_3)_2(aq)$

- ***Important Note: Calcium hydrogen carbonate $Ca(HCO_3)_2$ is always aqueous never solid***

Topic 5-Organic Chemistry

Introduction

Organic chemistry is the chemistry of carbon and its compounds except oxides, carbonates and hydrogen carbonates (bicarbonates). There are more compounds of carbon that all other elements put together.

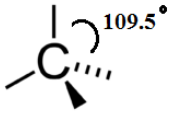
Why are compounds of carbon so special that a branch of chemistry was dedicated for them?

Bonding and Shape:

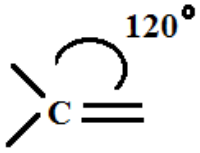
Covalent bonding that carbon atoms form have three unique features:

1. C-atoms can link to each other by strong covalent bonds forming very long chains.
2. C-atoms can link to each other by strong covalent bonds forming rings.
3. Every C-atom in an organic molecule must form 4 covalent bonds. These can be:

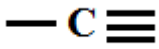
- 4 single bonds, which will be directed at the corners of a tetrahedral:



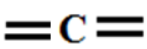
- 2 single bonds and 1 double bond, which will be directed at the corners of a triangle:



- 1 single bond and 1 triple bond, which will be directed at 180° from each other.



- 2 double bonds, which will be directed at 180° from each other.

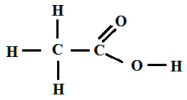
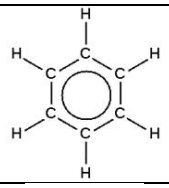
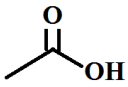


Organic compounds in which all bonds are single are called saturated, while those in which at least one double or triple bond are found are called unsaturated.

Organic molecules are usually represented by an empirical formula, a molecular formula, a structural formula, a displayed formula or a skeletal formula.

- empirical formula: shows the elements in their smallest whole number ratio
- molecular formula: shows the elements in their actual ratio
- displayed formula: shows every single element and every single bond
- structural formula: an unambiguous description of the way the atoms are arranged in a molecule
- skeletal formula: shows all atoms except C and H that are bonded together

For eg. consider acetic acid:

Empirical Formula	CH_2O	CH
Molecular formula	$C_2H_4O_2$	C_6H_6
Structural formula	CH_3COOH	C_6H_6
Displayed formula		
Skeletal formula		

The names of organic compound are divided into 2 parts:

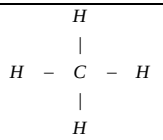
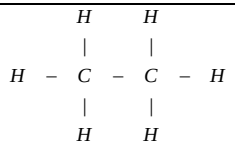
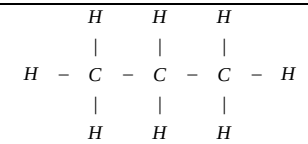
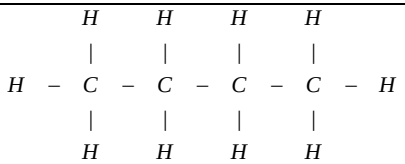
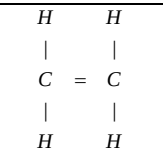
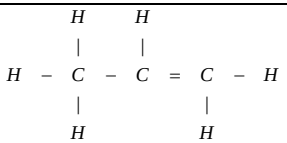
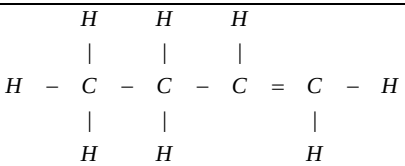
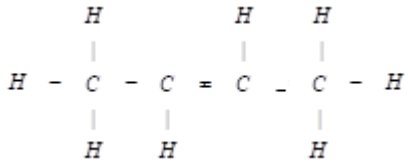
Prefix : number of carbon atoms
Suffix: functional group or class

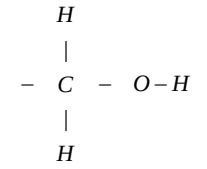
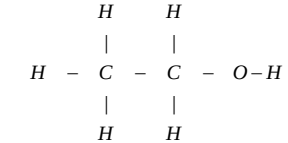
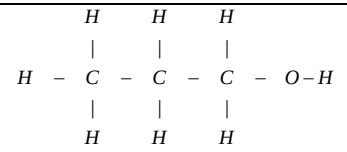
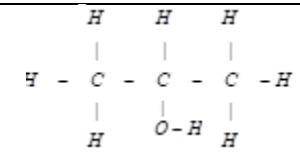
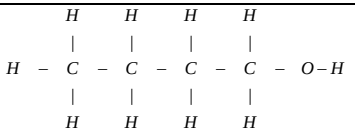
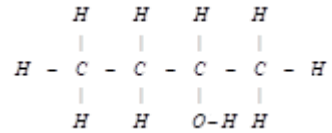
Prefix		Suffix	
Number of C-atoms	Name	Class	Name
1	*Meth-*	Alkanes	*-ane*
2	*Eth-*	Alkenes	*-ene*
3	*Prop-*	Alkanols	*-anol or -yl alcohol*
4	*But-*	Alkanoic acid	*-anoic acid*
5	*Pent-*	Ester	*-yl -anoate*
6	*Hex-*		

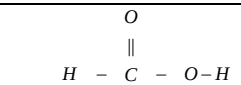
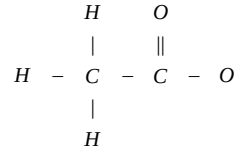
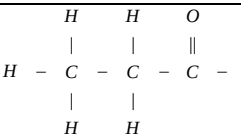
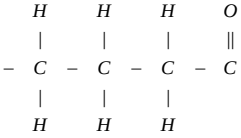
The simplest organic compounds are called alkanes. They are composed of carbon and hydrogen atoms ONLY, all bonds being single. Other organic families are derived from alkanes by replacing at least one H-atom by another atom or groups of atoms (called functional group) OR by replacing at least one of the single bonds with a double or triple one.

Alkane $\xrightarrow{-H}$ Alkyl group

For example: $CH_4 \xrightarrow{-H}$ CH_3-
Methane methyl-
$C_2H_6 \xrightarrow{-H}$ C_2H_5-
Ethane ethyl-

Family	Example	Molecular formula	Structural formula	Displayed formula	Boiling point (°C)	State at RTP
Alkane C_nH_{2n+2}	Methane	CH_4	CH_4		- 162	Gas
	Ethane	C_2H_6	CH_3CH_3		- 89	Gas
	Propane	C_3H_8	$CH_3CH_2CH_3$		- 42	Gas
	Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$		0	Gas
Alkene C_nH_{2n}	Ethene	C_2H_4	$CH_2=CH_2$		- 104	Gas
	Propene 1-Propene Prop-1-ene	C_3H_6	$CH_2=CHCH_3$		- 48	Gas
	1-Butene But-1-ene	C_4H_8	$CH_2=CHCH_2CH_3$		- 6	Gas
	2-Butene But-2-ene	C_4H_8	$CH_3CH=CHCH_3$		3.7	Gas

Family	Example	Molecular formula	Structural formula	Displayed formula	Boiling point (°C)	State at RTP
Alcohol $C_nH_{2n+1}OH$	Methanol Or Methyl alcohol	CH_3OH	CH_3OH		65	Liquid
	Ethanol Or Ethyl alcohol	C_2H_5OH	CH_3CH_2OH		76	Liquid
	1-Propanol Or Propan-1-ol	C_3H_7OH	$CH_3CH_2CH_2OH$		97	Liquid
	2-Propanol Or Propan-2-ol	C_3H_7OH	$CH_3CH(OH)CH_3$		82	Liquid
	1-Butanol Or Butan-1-ol	C_4H_9OH	$CH_3CH_2CH_2CH_2OH$		118	Liquid
	2-Butanol Or Butan-2-ol	C_4H_9OH	$CH_3CH(OH)CH_2CH_3$		99.5	Liquid

Family	Example	Molecular formula	Structural formula	Displayed formula	Boiling point (°C)	State at RTP
Acids $C_nH_{2n+1}COOH$	Methanoic acid Or Formic acid	HCO_2H	HCOOH		101	Liquid
	Ethanoic acid Or Acetic acid	CH_3CO_2H	CH_3COOH		118	Liquid
	Propanoic acid	$C_2H_5CO_2H$	CH_3CH_2COOH		141	Liquid
	Butanoic acid	$C_3H_7CO_2H$	$CH_3CH_2CH_2COOH$		164	Liquid

Hydrocarbons

Definition of hydrocarbons
Hydrocarbons are organic compounds made up of carbon and hydrogen **only**.

CQD4. [G] What are hydrocarbons? Which of the following compounds are classified as hydrocarbons?

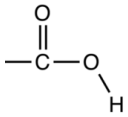
a.C_3H_8 b.C_3H_7OH c. CH_3COOH d. C_6H_6

The simplest hydrocarbon families are:

a) Alkanes, which are **saturated** (all carbon-carbon bonds are single) unreactive hydrocarbons.
b) Alkenes which are **unsaturated** (at least one C = C bond) reactive hydrocarbons
c) Alkynes which are **unsaturated** (at least one C ≡ C bond) reactive hydrocarbons

Homologous series: A family of similar compounds with similar chemical properties due to the presence of the same functional group.

Functional Group: An atom or group of atoms that determine the chemical properties of a homologous series.

Family	Functional Group
Alkane C_nH_{2n+2}	**None**
Alkene C_nH_{2n}	**C = C**
Alcohol $C_nH_{2n+1}OH$	**-OH**
Acids $C_nH_{2n+1}COOH$	**-COOH** 

Members of the same homologous series:

- Have the same general formula.
- Have the similar chemical properties.
- Have the same functional group.
- Differ from the one before and the one after in the series by a CH_2 unit.
- Show a trend in physical properties, e.g. melting point and boiling point will generally increase with increasing carbon chain length.

CQD1. a. Define the term homologous series.
b. List three general properties of members in the same homologous series.

Alkanes

- General formula: C_nH_{2n+2}
- **Saturated** hydrocarbon: Every carbon atom forms 4 single covalent bonds (directed at the corners of a tetrahedral) with four other atoms. No new atoms can be added to an alkane molecule.
- It is an unreactive hydrocarbon.
- Has no specific functional group.

CQD5. [G]
a. Give the general molecular formula for alkanes.
b. Give the molecular formula for an alkane having three carbons.

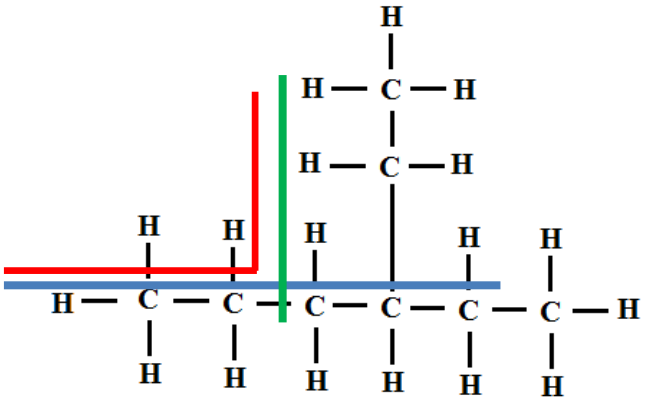
Nomenclature:

Unbranched:

- Use a Latin prefix to indicate the number of carbons.
- Use the suffix -ane.

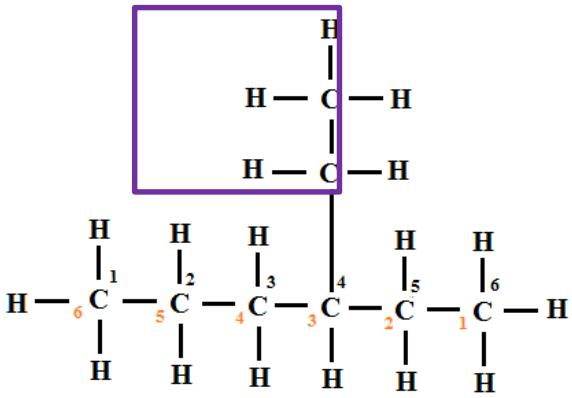
Branched:

- Choose the longest chain you can find, count the number of carbon atoms.



Both the red and blue contain 6 carbons, choose either.

- Number the carbons in your chosen chain from whichever end gives the carbon carrying the branch the smallest number.

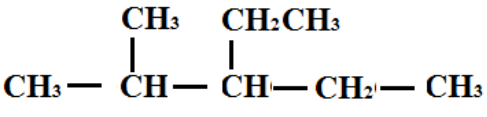


The carbon carrying the branch is numbered either 3 or 4, choose 3.

- Name the branch. Branches of the formula C_nH_{2n+1} are called alkyl groups. They are named using the prefixes and the ending –yl. The branch above has 2C-atoms ⇒ we use the prefix eth and ending –yl ⇒ ethyl.

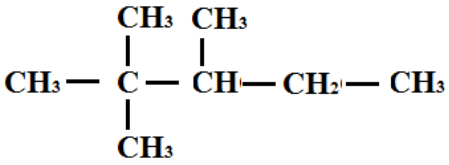
Now we are ready to name, the longest chain of C contains 6 carbons ⇒ hex-
It is an alkane ⇒ hexane
It has an ethyl branch ⇒ ethylhexane
The branch is on carbon 3 ⇒ 3-ethylhexane

- If the parent chain contains more than one branch, we must indicate the position of each and list them alphabetically. For eg.



is called 3-ethyl-2-methypentane

- If the same branch exists at different carbons, we name it once using di, tri, … to indicate how many. For eg.



is called 2,2,3-trimethypentane

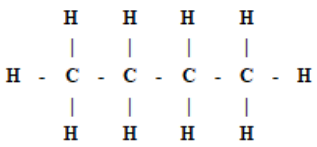
- The branches may be groups other than alkyl, the same rules apply. For example:

  $CH_3CHBr\text{-}CH_2Br$ is called 1, 2-dibromopropane.

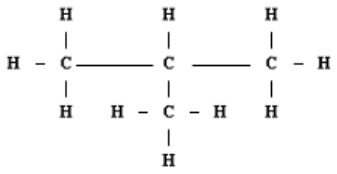
Structural isomers: They are organic compounds having the same molecular formula but different structural formulae.

Isomers of Alkanes
Alkanes with more than 4 carbons have isomers. There are two isomers of the formula C_4H_{10}:

n-butane



2-methylpropane



CQD7. [G] Hexane, C_6H_{14}, has five structural isomers. Draw the condensed structural formulas for isomers of hexane and give the name for each.

Melting and boiling Points:

- As the number of C-atoms in straight chain alkanes increases, their melting points, boiling points and densities increase. This is because the strength of the weak forces between the molecules (intermolecular forces) increases. This is observed to be truein all homologues series.
- Branching decreases boiling points but increases melting points

Reactions:

Generally, alkanes are inert, they do however undergo combustions, cracking and chlorination.

Combustion

Alkanes are often used as fuels. Methane is the major component of natural gas and compressed natural gas (CNG) has been used as an alternative to gasoline for vehicles. LPG (liquified petroleum gas) is a mixture of propane and butane, used also in vehicles. Cylinders of butane gas are used in homes, camping stoves, blowtorches and gas lighters.

The reaction when a hydrocarbon is burnt in a sufficient amount of oxygen for all its carbons to become carbon dioxide $CO_2(g)$ and all its hydrogens to become water is called **complete combustion**.

For example;

Word equation: methane + oxygen → carbon dioxide + water

Formula equation: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

In the presence of insufficient amount of oxygen, carbon monoxide CO(g) or C(s) and water are produced. This is known as **incomplete combustion**.

Word equation: methane + oxygen → carbon monoxide + water

Formula equation: $2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$

The appearance of the flame resulting from burning an organic compounds gives us clues about its nature. Organic compounds with a low % of C-atoms usually burn with a clean blue flame, while compounds with a high % of C-atoms usually burn with a smoky (sooty) orange flame. The fine particles of carbon resulting from incomplete combustion glow yellow in the heat of a flame and impart the characteristic yellow color to it.

CQD2. Write an equation to represent the complete combustion of C3H8 (ignore state symbols in your equation).

$$\mathbf{2C_3H_8 + 10O_2 \rightarrow 6CO_2 + 8H_2O}$$

Disadvantages of incomplete combustion:

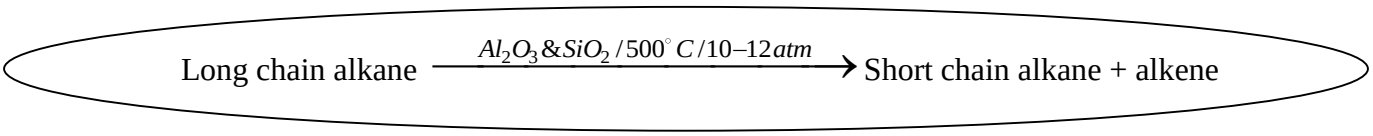
Incomplete combustion is wasteful and uneconomic. The full energy content of the fuel does not get released. Moreover, the formation of carbon or soot reduces the efficiency of the burner or engine being used.

Furthermore, formation of carbon monoxide is dangerous as it is extremely poisonous. Hemoglobin, the molecules that carry oxygen around in our bodies, has a greater affinity for CO than oxygen. In other words, in the presence of CO, hemoglobin molecules attach or carry CO rather than oxygen. Brain cells deprived of their supply of oxygen will quickly die.

Cracking

It is the process of breaking down long chain alkanes (less useful) to shorter alkanes (more useful) and alkenes. This can be done by c**atalytic cracking**, the decomposition of long-chain alkanes into alkenes and smaller alkanes using a catalyst and high temperature. Typical conditions are:

- Al_2O_3 & SiO_2 / 500°C/ 10-12 atm, and
- Cr_2O_3/ 500°C/ 10-12 atm



Example

$$C_8H_{18} \xrightarrow{Al_2O_3\&SiO_2\,/\,500^\circ C\,/\,10\text{–}12\,atm} C_2H_4 + C_3H_6 + C_3H_8$$

Octane — ethene + propene + propane

In all cracking reactions the products will include a shorter chain alkane and a shorter chain alkene or two or more alkenes and hydrogen.

➢ Applications:

Q : Complete the following reactions (give formula and name)

a) $C_7H_{16} \xrightarrow{Al_2O_3\&SiO_2\,/\,500^\circ C\,/\,10\text{–}12\,atm} 2$ ……. + C_3H_8

heptane — propane

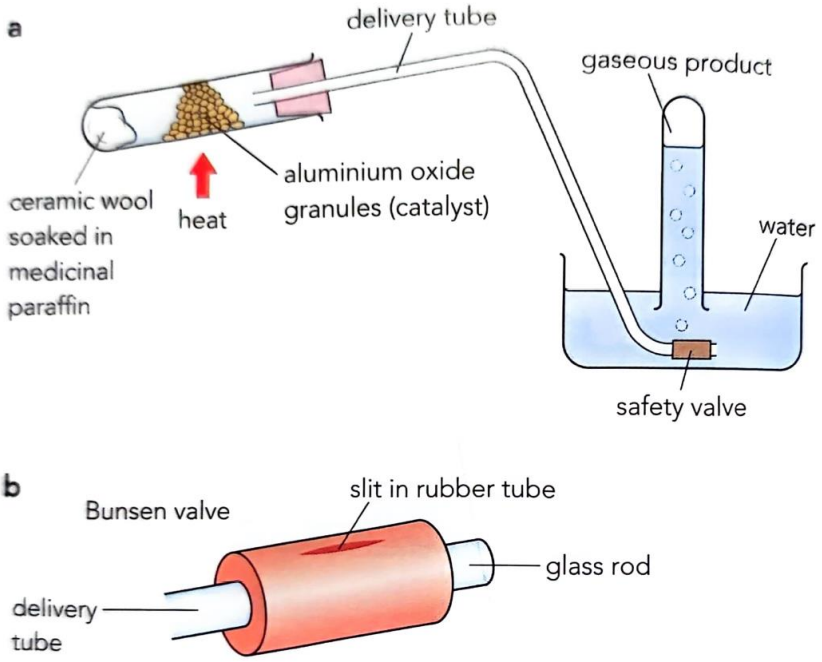
b) $C_7H_{16} \xrightarrow{Al_2O_3\&SiO_2\,/\,500^\circ C\,/\,10\text{–}12\,atm} C_2H_4$ + …….

heptane — ethene

c) $C_{10}H_{22} \xrightarrow{Al_2O_3\&SiO_2\,/\,500^\circ C\,/\,10\text{–}12\,atm}$ …… + C_2H_4

decane — ethene

The figure below shows a simple apparatus that can be used to carry out cracking reactions in the laboratory. You will notice that in the laboratory we may use a catalyst of broken, unglazed pottery.



From Cambridge IGSCE Chemistry Textbook

Chlorination

Chlorination is a substitution reaction.

A substitution reaction is a reaction in which one atom or a group of atoms is (are) replaced by another atom or group of atoms, without changing the molecule's general structure.

Chlorination is a photochemical reaction, ie. a reaction where the activation energy required to start the reaction is provided by light, usually of a particular wavelength, shining of reactants.

In the case of chlorination, sunlight, or any other source of ultraviolet light, provides the activation energy needed to split the chlorine molecules into atoms. When this type of reaction takes place, these atoms are called **free radicals** and they are very reactive.

When methane reacts with chlorine, the products of the reaction depend on whether there is an excess of methane or an excess of chlorine. If there is an excess of methane it forms chloromethane and hydrogen chloride.

$$\text{Methane} + \text{Chlorine} \xrightarrow{U.V} \text{Chloromethane} + \text{Hydrogen chloride}$$

$$CH_4 + Cl_2 \xrightarrow{U.V} CH_3Cl + HCl$$

Because we cannot control the chlorine free radicals produced in this reaction, we also obtain small amounts of other 'substituted' products CH_2Cl_2(dichloromethane), $CHCl_3$ (trichloromethane or chloroform) and CCl_4 (tetrachloromethane or carbon tetrachloride).

This reaction will NOT occur in the dark.

CQD8. [G] Methane, CH_4, undergoes a substitution reaction with chlorine, Cl_2.

a. What is the factor that speeds up this reaction?
b. Give the chemical equation for the substitution reaction between a methane molecule and a chlorine molecule.

Alkenes

- General formula: C_nH_{2n}
- Unsaturated hydrocarbon (contain a carbon-carbon double bond)
- It is a reactive hydrocarbon.
- Functional group: carbon-carbon double bond (C=C)

CQD6. [G]
a. Give the general molecular formula for alkenes.
b. Give the molecular formula of an alkene having two carbons.

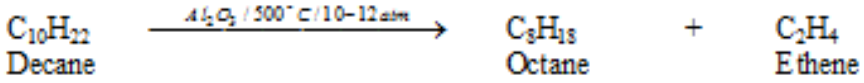
CQD10. [G] Classify the following into saturated and unsaturated hydrocarbons.

a. 2-heptene, C_7H_{14}
b. hexane, C_6H_{14}
c. C_5H_{12}
d. methylpropane, C_4H_{10}
e. cyclopentane, C_5H_{10}

Preparation of Alkenes

As discussed earlier, long chain alkanes are not very useful and in low demand compared to short chain alkanes and alkenes. So, they are usually cracked to produce the more in demand hydrocarbons.

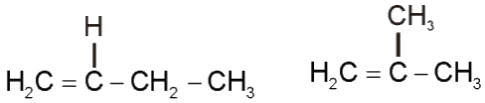
Industrially, alkenes are obtained by cracking of long chain alkanes using a catalyst and high temperature.
For example;



Isomers of Alkenes

- Isomers can be the result of the functional group, for eg. C_4H_8:

- Or chain length like alkanes:



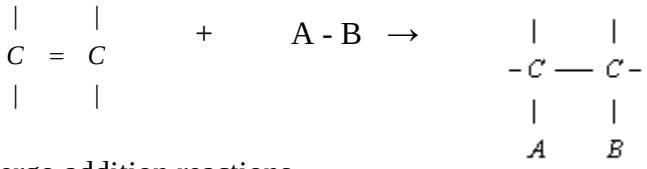
Reactions

Combustion

Example: $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O + \text{heat}$

Addition Reactions:

An addition reaction is one in which 2 or more molecules combine to form **one** molecule.



Being unsaturated, alkenes easily undergo addition reactions.

a) Addition of hydrogen (hydrogenation)

Conditions: 150-300°C temperature in the presence of nickel catalyst

$$\text{Alkene + Hydrogen} \xrightarrow{Ni/200^{\circ}C} \text{Alkane}$$

Example:

$$\text{Ethene} + H_2(g) \xrightarrow{Ni/200^{\circ}C} \text{Ethane}$$

$$C_2H_4 + H_2(g) \xrightarrow{conc\ H_2SO_4/80^{\circ}C} C_2H_6$$

b) Addition of steam (hydration)

Conditions: Concentrated sulphuric acid at 80°C **OR** Phosphoric acid as a catalyst at 300°C and 70 atm

$$\text{Alkene + Steam} \xrightarrow{Conditions} \text{Alcohol}$$

Example:

$$\text{Ethene} + \text{Steam} \xrightarrow{conc\ H_2SO_4/80^{\circ}C} \text{Ethanol}$$

$$C_2H_4 + H_2O \xrightarrow{conc\ H_2SO_4/80^{\circ}C} CH_3CH_2OH$$

c) Addition of bromine (bromination)

Alkene is mixed with bromine dissolved in an organic solvent like CCl_4.

$$\text{Alkene + Bromine} \xrightarrow{Room\ Conditions} \text{1,2-dibromoalkane}$$

Example:

$$\text{Ethene} + \text{Bromine} \rightarrow \text{1,2-dibromoethane}$$

$$C_2H_4 + Br_2(CCl_4) \rightarrow CH_2BrCH_2Br$$

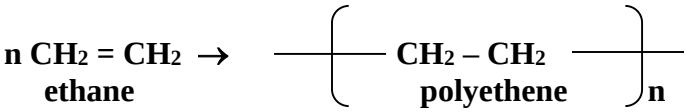
If the alkene is mixed with bromine dissolved in water, a different product will be formed.

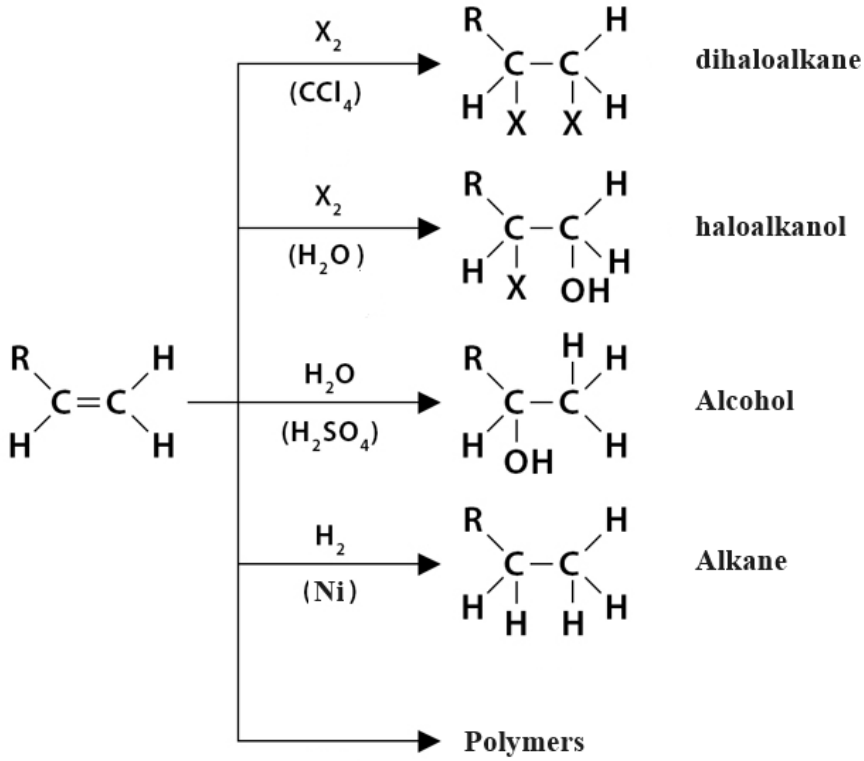
$$C_2H_4 + Br_2(aq) \rightarrow CH_2BrCH_2OH$$

This reaction is used to differentiate between saturated and unsaturated hydrocarbons.
Mixing unsaturated hydrocarbons with brown bromine water (or $Br_2(CCl_4)$ rapidly decolourizes it in the dark.
Alkanes do not react with $Br_2(aq)$ in the dark.

d) Polymerization

Alkene molecules tend to add to each other forming very large molecules (like a pearl necklace). Such reactions in which many small molecules (monomers) from a single very large molecule (polymer) are called polymerization reactions.





Note that with longer chain alkenes, more than one product may be formed when alkenes react. For example,

$CH_3CH_2CH{=}CH_2 + H_2O \rightarrow CH_3CH_2CH(OH)CH_3 + CH_3CH_2CH_2CH_2OH$

CQD3.

a. Is the ethene molecule planar or 3-dimensional?

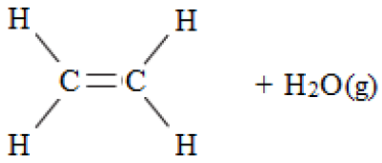
b. Write an equation to show how ethene is produced from ethanol (ignore state symbols in your equation). What conditions are needed for the reaction to occur?

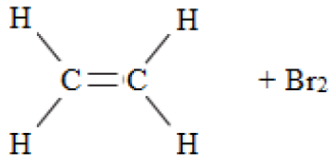
c. Write an equation for the complete combustion of C_3H_6 (ignore state symbols in your equation). Explain why the flame produced is colored, unlike that of C_3H_8.

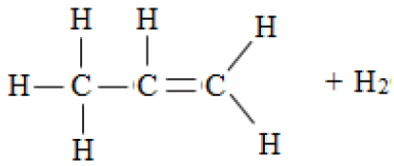
d. Write equations showing what happens when C_2H_4 reacts separately with hydrogen and hydrogen chloride (ignore state symbols in your equation). Give the name of the product in each case. What name is given to this type of reaction and why?

e. How can you distinguish between ethane, C_2H_6, and ethene C_2H_4? Write equations for any possible reaction(s).

CQD9. [G] Give the structural formula for the product of each of the following chemical reactions.

a. 

b. 

c. 

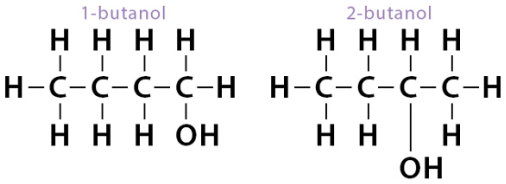
Alcohols

General formula: $C_nH_{2n+1}OH$ (or R—OH, where R represents an alkyl group)
Functional group: hydroxyl group (-OH)

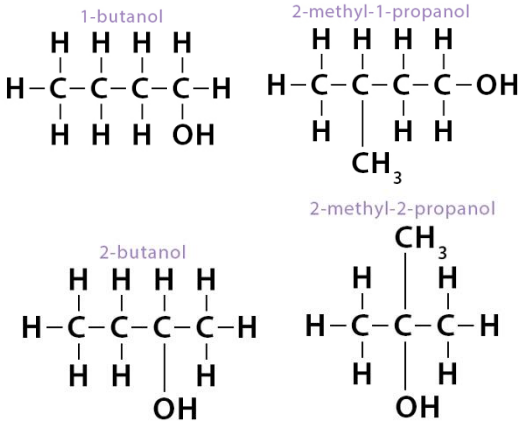
CQD12. [G]
a. Give the general molecular formula for alkanols (alcohols).
b. Give the molecular of an alkanol having two carbons.

Isomers:
Like alkenes, isomers of alcohols can be due to the position of the functional group as well as the chaon length.

- Isomers can be the result of the functional group, for eg. C_4H_9OH.



- Or chain length like alkanes:



Preparation of Ethanol:

a. Fermentation

Fermentation: anaerobic respiration of yeasts (i.e absence of oxygen or no oxygen) at a temperature between 25°C and 35°C.

$$\text{glucose} \xrightarrow{yeast} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6 \xrightarrow{yeast} 2C_2H_5OH + 2CO_2$$

Conditions of fermentation:

1. Yeast, a microorganism that produces enzymes (biological catalysts)
2. Optimum temperature of 25 - 35°C
3. Absence of oxygen or anaerobic conditions (Air-lock system)
   To prevent oxidation of ethanol to ethanoic acid or vinegar that has a sour taste.

Note:
Fermentation is self-limiting because ethanol is toxic to yeasts. When the concentration of ethanol in the mixture reaches around 14% (or when the sugar runs out), yeasts die and fermentation stops.

At the end of fermentation process, pure alcohol is extracted by fractional distillation.

B. Hydration of ethene (Industrial method)
Ethene is obtained from refining of crude oil and then mixed with steam and compressed to 60 atm (6000kPa), passed over a catalyst of immobilized phosphoric(V) acid at 300°C. high purity ethanol is produced.

$$\text{Ethene} + \text{Steam} \rightarrow \text{Ethanol}$$

Conditions: Phosphoric acid as a catalyst at 300°C and 6000kPa/60 atm

$$\text{Ethene} + \text{Steam} \xrightarrow{H_3PO_4/300^\circ C/60\ \text{atm}} \text{Ethanol}$$

$$C_2H_4(g) + H_2O(g) \xrightarrow{H_3PO_4/300^\circ C/60\ \text{atm}} C_2H_5OH$$

Comparing the two methods of preparing ethanol

	Fermentation (Laboratory method)	**Hydration of alkenes (Industrial method)**
Resources	Uses renewable resources (sugar cane, sugar beet)	Uses non-renewable resources (petroleum)
Type of process	A batch process Everything is mixed together in a reaction container and then left for several days. The batch is then removed and a new reaction is set up again. This is inefficient.	A continuous process A stream of reactants is constantly passed over the catalyst. This is a more efficient than a batch process.
Rate of reaction	Slow, taking several days for every batch	Rapid
Equipment	Small-scale equipment capable of withstanding pressure	Relatively simple, large vessel
Quality of product	Produces very impure ethanol that needs further processing (fractional distillation)	Produces much purer ethanol
Reaction conditions	Uses gentle temperatures and ordinary pressure Simple, straightforward method	Uses high temperatures and pressures, needing a high input of energy. Sophisticated, complex method
Cost	Lower	Higher

The chosen method depends on many factors. Examples:

1. Availability of resources: If the country does not have abundant crude oil but the climate conditions and the type of land needed for growing sugar cane is available, then economically, fermentation is used to produce ethanol rather than direct hydration.
2. If pure ethanol is required, to be used as a solvent, then economically, the continuous process of direct hydration is used.
3. For the production of wine, only dilute ethanol is needed, so in this case fermentation is a good choice.

Uses of Ethanol

Use	Reason
As a fuel (for cars).	It only releases CO_2 and H_2O into the environment, no other pollutant gases like petroleum. It is a renewable source because it comes from plants as sugar beet or sugar cane
As a solvent in (in paints, perfumes, glues, aftershaves, inks…)	due to the presence of –OH group that allows the alcohol to dissolve in water and to dissolve other organic compounds It is volatile
In beverages and methylated spirits	In ancient times, it was believed that fermented drinks were safer than polluted water. Alcohols affect the brain, it causes depression, slows reactions and human responses. Alcohol is poisonous and damages the liver
raw material to make other compounds	Highly reactive

CQD13. [G] Ethanol can be prepared biologically from glucose.

a. Name the biological process of preparation of ethanol from glucose.
b. Give the chemical equation for this process.
c. What are the conditions needed for this reaction to occur?
d. Suggest a technique to separate ethanol from final solution.

Reactions

a) Combustion

Ethanol is known as a biofuel (a fuel made from biological sources such as sugar cane).
The introduction of the use of ethanol is increasing nowadays to reduce the use of fossil fuels that are non-renewable sources.

$$\text{Alcohol} + \text{Oxygen} \rightarrow \text{Carbon dioxide} + \text{Water}$$

Example:

$$\text{Ethanol} + \text{Oxygen} \rightarrow \text{Carbon dioxide} + \text{Water} + \text{heat}$$

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O + \text{heat}$$

Ethanol burns with a clean flame.

Q: The flame of which of the following tends to be the cleanest upon combustion: C_2H_4, C_2H_6 or C_2H_5OH?

b) Oxidation of alcohols

$$\text{Alcohol} \xrightarrow{[o]} \text{carboxylic acid} + \text{water}$$

Possible oxidizing agents:

a) Hot acidified potassium permanganate $\underline{KMnO_4 / H^+}$ (colour changes from purple to colourless)
b) Hot acidified potassium dichromate $\underline{K_2Cr_2O_7 / H^+}$ (colour changes from orange to green)

c) Esterification

$$\text{Acid} + \text{alcohol} \leftrightarrows \text{ester} + \text{water}$$

Acid catalyst needed

$$\text{Ethanol} + \text{acetic acid} \xrightleftharpoons{\text{Concentrated } H_2SO_4} \text{ethylethanoate} + \text{water}$$

$$C_2H_5OH + CH_3COOH \xrightleftharpoons{\text{Concentrated } H_2SO_4} C_2H_5OOCCH_3 + H_2O$$

CQD11.

a. Give another name for methanol and ethanol.
b. Ethanol is produced from sugar by fermentation.
Fermentation produces a dilute solution of alcohol. How can the product be concentrated?
c. Write an equation to show how ethanol is produced from ethene. What conditions are needed for the reaction to occur? (Ignore state symbols in your equation.)
d. List three common uses of ethanol.
g. What is produced when ethanol is warmed with a suitable oxidizing agent?

Carboxylic Acids

General formula: R-COOH or $C_nH_{2n+1}COOH$
Functional group: -COOH group (carboxylic acid group)

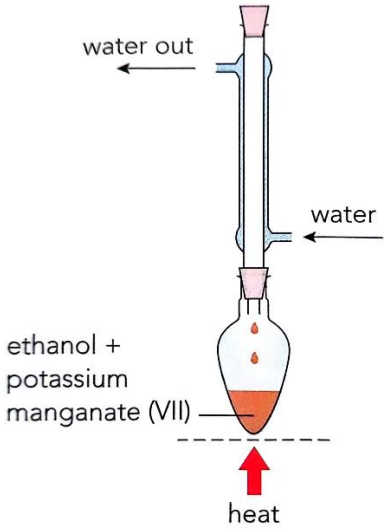
Preparation of Acids

Acids can be made by the oxidation of alcohols

- with warm acidified potassium manganate(VII), $KMnO_4$
- by bacterial oxidation during vinegar production

$$C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$$

The ethanol is mixed with the acidified potassium manganate(VII) solution and heated in a round bottomed flask under reflux. The condenser is set up in a vertical position to prevent theh volatile alcohol from escaping. Remember that alcohol is highly flammable and heating it should be done with extreme care and never with direct flame.



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Reactions of Acids-

a. As a weak acid

Ethanoic acid is a **typical weak acid** that partially dissociates or ionizes in water to give a low concentration of hydrogen ions.

Ethanoic acid ⇆ Ethanoate ion + hydrogen ion

$CH_3COOH(aq) \leftrightarrows CH_3COO^-(aq) + H^+(aq)$

1. Reaction with metals

ethanoic acid + magnesium → magnesium ethanoate + hydrogen

$2CH_3COOH + Mg \rightarrow Mg(CH_3COO)_2 + H_2$

magnesium ethanoate

Rule for naming of the organic acid salt: name of metal + Latin prefix for # of C-atoms + suffix -oate

2. Reaction with metal hydroxides

ethanoic acid + sodium hydroxide → sodium ethanoate + water

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

3. Reaction with metal carbonates

ethanoic acid + calcium carbonate → calcium ethanoate + Water + Carbon dioxide

$2CH_3COOH + CaCO_3 \rightarrow Ca(CH_3COO)_2 + H_2O + CO_2$

CQD14.

a. Name and give the molecular and structural formulas of the first three organic acids.

b. What methods can be used to produce ethanoic acid?

c. Write an equation to show how ethanoic acid can be produced in the lab. What conditions are needed for the reaction to occur?

d. What gas is produced when ethanoic acid reacts with zinc, Zn?

CQD15. [G] Ethanoic acid, CH_3COOH, can be prepared from ethanol, C_2H_5OH, or ethanal, C_2H_4O.

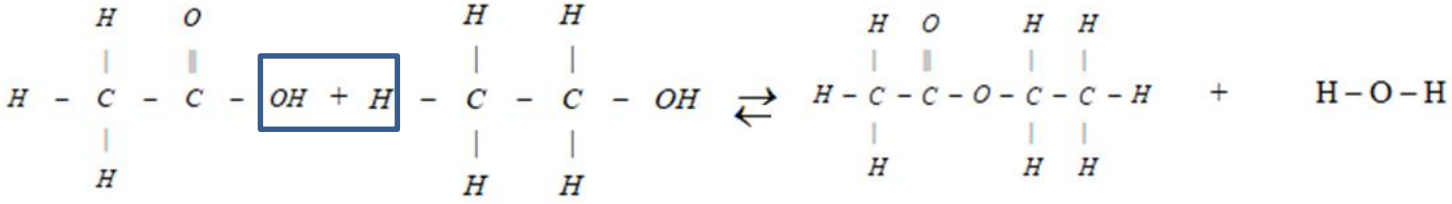
a. Identify the type of such reaction.

b. Which of the following reagents are needed for preparation of ethanoic acid from ethanol or ethanal?

- acidic dichromate solution
- acidic permanganate solution
- concentrated solution of sulfuric acid
- concentrated solution of sodium hydroxide
- steam at high pressure

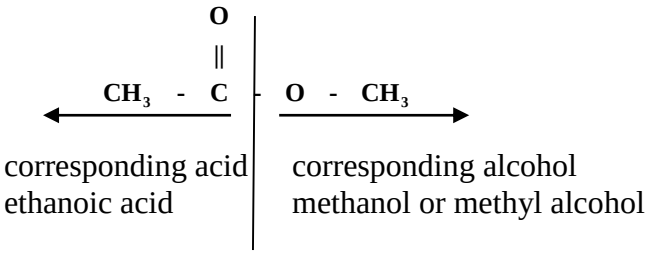
b. **Esterification**

Alcohols react with acids to produce the sweet-smelling esters. The reaction requires an acid catalyst, usually concentrated sulfuric acid, and sets up an equilibrium.



To name an ester, use

the alkyl group of alcohol + Latin prefix for number of carbon atoms of acid + suffix –oate



Example:

CH_3COOCH_3	Ester:	methyl ethanoate
	Alcohol:	methanol
	Acid:	ethanoic acid

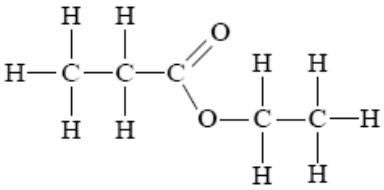
Different combinations of alcohols and acids can produce esters that are isomers. For example, ethanol with propanoic acid and propanol with ethanoic acids.

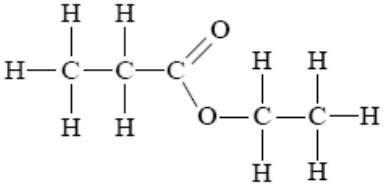
$CH_3CH_2COOH + C_2H_5OH \rightleftarrows CH_3CH_2COOCH_2CH_3 + H_2O$

Ethyl propanoate

$CH_3COOH + C_3H_7OH \rightleftarrows CH_3COOCH_2CH_2CH_3 + H_2O$

Propyl ethanoate

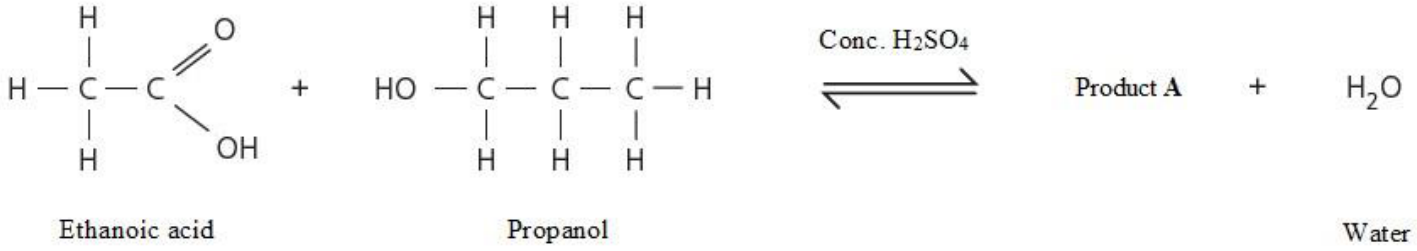




CQD16.

a. Draw the general structural formula of an ester and circle the functional group.
b. List two common uses of esters.
c. Which ester will be produced if ethanol and methanoic acid react?

CQD17. [G] The following equation represents the reaction between ethanoic acid and propanol.



a. Give the name of this reaction.
b. Draw the structural formula for product **A**.
c. Give the name of product **A**.
d. Compound **B** has the same molecular formula of compound **A**. Choose the name of the carboxylic acid and the alcohol that react together to form compound **B**.

	Carboxylic acid	**Alcohol**
A	methanoic acid	propanol
B	butanoic acid	ethanol
C	propanoic acid	ethanol
D	ethanoic acid	methanol

Fuels

Fossil fuels are found in the Earth's crust and were formed millions of years ago by the action of high pressure, high temperature and bacteria on once living material. The three main fossil fuels are petroleum, natural gas and coal.

- crude oil or petroleum is a black, thick, sticky liquid found under the sea and under the ground in certain parts of the world. It is a mixture of many different hydrocarbons.
- methane is the main constituent of natural gas (94%).

Crude oil and natural gas were formed from the bodies of marine organisms.

- coal is a black rock formed from prehistoric plant remains, composed largely of carbon.

Fossil fuels are non-renewable fuels (finite resource) since a very long time is needed for their formation and so they must be used efficiently.

CQD18. Petroleum and coke are considered fossil fuels. Are they renewable sources of energy? Explain your answer.

Petroleum refining

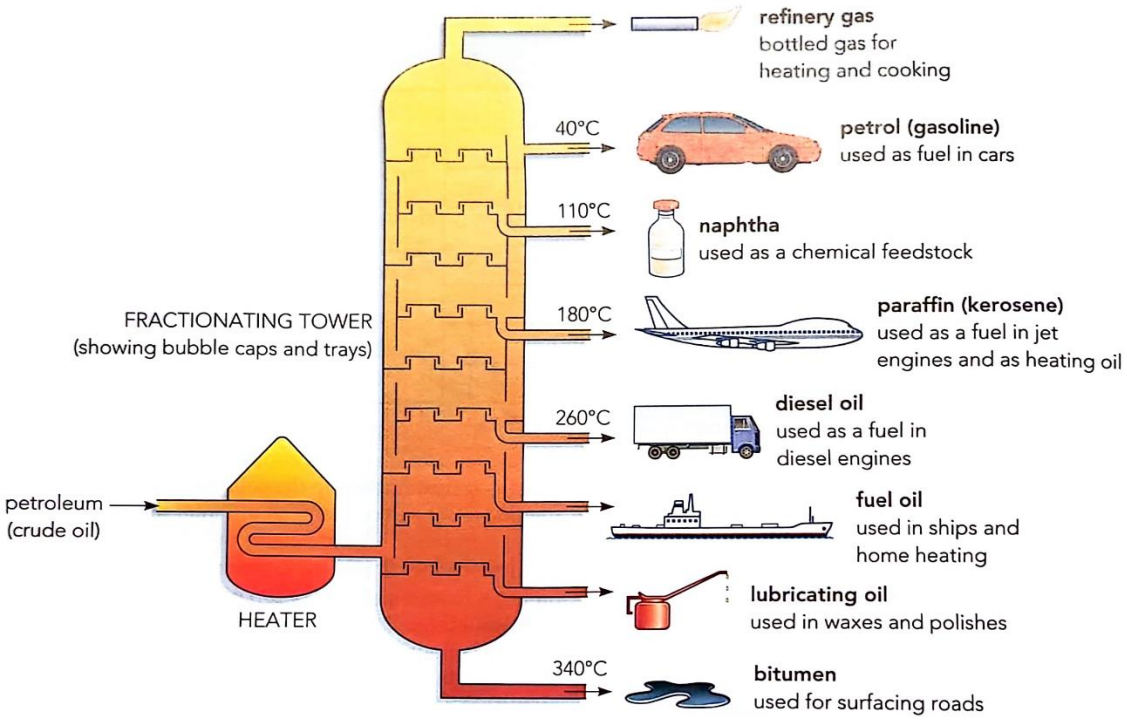
Around 90% of the petroleum extracted is used to make fuels and the rest is used as chemical feedstock (chemical used as a raw material in industry to make useful chemical products).

Crude oil cannot be used as is, that is why it has to be first refined.
Refining is the process of separating petroleum into useful products, this is done by two processes:

1. Fractional distillation of petroleum.
2. Purification of the fractions.

Separating the fractions of crude oil

Petroleum is strongly heated first to convert it to vapor. The vapor is fed into the bottom of a **fractionating column**. The hydrocarbons that have very high boiling points turn immediately into liquids and are run off at the bottom of the column. The different fractions will condense at different heights according to their different boiling points. When they condense they are removed as liquids. Some will not condense at all and will leave the column as gases. The fractions that come out near the top of the column are light-coloured runny liquids. The fractions collected at the bottom are thick or **viscous**.



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The properties of petroleum's different fractions change from the top to the bottom of the fractionating column. As we go down the column, the fractions':

- chain length increases
- volatility decreases
- boiling points increase
- viscosity increases

Summary of the different petroleum fractions and their uses:

Fraction	Use
Refinery gas (a mixture of methane, ethane, propane and butane, for bottled gas)	For heating and cooking
Gasoline or Petrol	Fuel in cars
Naphtha	Chemical feedstock
Kerosene or Paraffin	Jet fuel
Diesel oil or Gas oil	Fuel in diesel engines (as tractors and lorries)
Fuel oil	Fuel for ships and heating systems
Lubricants and waxes	Lubricants, waxes and polishes
Bitumen (tar)	Making roads

Cracking:

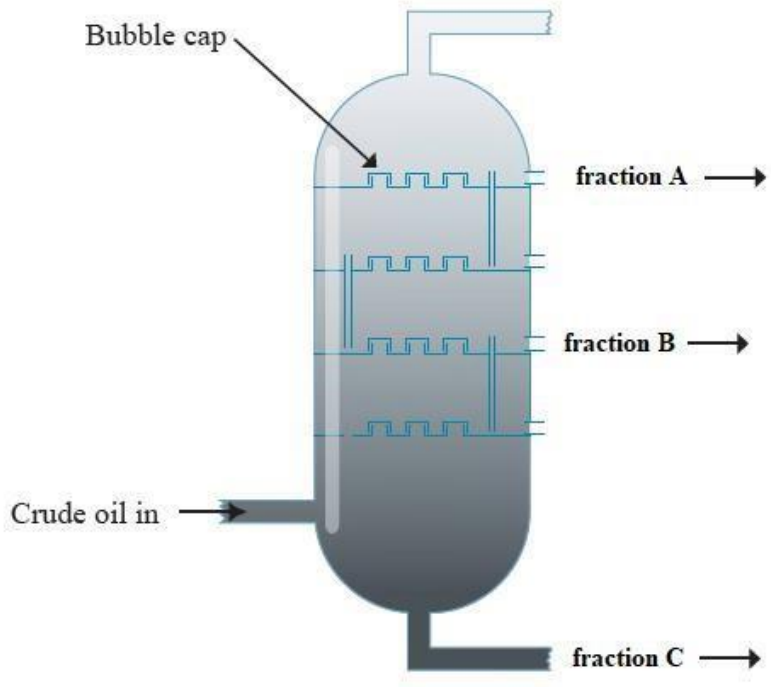
A large proportion of the products of fractional distillation are long chain hydrocarbons (heavy oils and bitumen). Remember that these are of limited use and are often broken down into shorter chain hydrocarbons which are more useful by cracking. There are 2 types of cracking:

1. **Thermal cracking**: temperature 700˚C and pressure of 30 atm
2. **Catalytic cracking**: temperature 500˚C, normal atmospheric pressure and the catalyst: mixture of silicon (IV) oxide (silica) and aluminium oxide (alumina)

CQD19. List four fractions obtained from the fractional distillation of crude oil, and state one common use of each.

CQD20. Distillation normally produces large amounts of long, straight-chain hydrocarbons, which have limited usage. On the other hand, short branched-chain hydrocarbons, which are in high demands, are normally produced in relatively low quantities. What process is used to convert the less useful fractions into more useful fractions?

CQD21. [G] The following diagram shows the tower used to separate petroleum into fractions.



a. Give the name of the process used to separate petroleum into its components.
b. Which physical property is the base for this process?
c. Which is the most volatile fraction, A, B, or C?

Macromolecules or Polymers

We can find polymers in all living things, most prominent examples being proteins and carbohydrates. Polymers have invaded our everyday life, most prominent examples being plastics.

Polymers are large molecules made up of smaller molecules called monomers joining together. "monomers" (mono = one, mer = unit) and "polymers"(poly = many).

Polymerization:
It is the process of linking together relatively small and simple molecules to form large macromolecular units called polymers.

Types of Polymers

Addition Polymers	Condensation polymers
A single product formed $n\,M \xrightarrow{\text{catalyst/high pressure}} (M)_n$ monomer polymer M = monomer n = number of monomers	It is the joining of molecules with the removal of a small molecule such as water or hydrogen chloride.
Single type of monomer only homopolymers	Monomers could be different or similar Copolymers or homopolymers
The monomer must be unsaturated	Monomers must have 2 functional groups which could be similar or different.

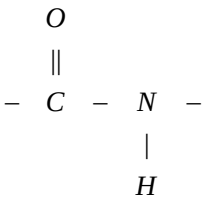
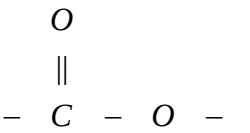
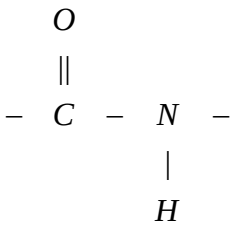
Polymers can also be classified into natural and synthetic. Natural being the ones found in nature and synthetic being the manmade.

CQD22.
a. Define the following terms:
i. Polymerization
ii. Monomer
b. Name the two classes of polymers. Name the two types of polymerization processes.

CQD26.[G] Polymers can be either synthetic or natural.
What are polymers? How are they formed?

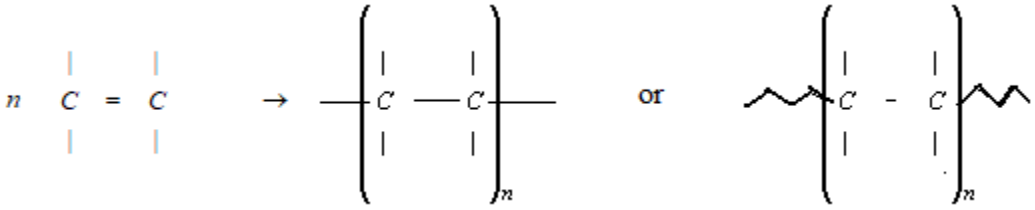
Below a summary of some common polymers you need to know.

Classes of polymers

Natural Polymer	**Synthetic polymers**		
condensation polymer	Addition polymers	Condensation polymers	
		Polyesters	Polyamides
Protein (**Polyamide)** Peptide or amide bond 	Polyethene Polypropene Polytetrafluorthene Polystyrene PVC	Terylene Or Dacron Ester bond 	Nylon Peptide or amide bond 

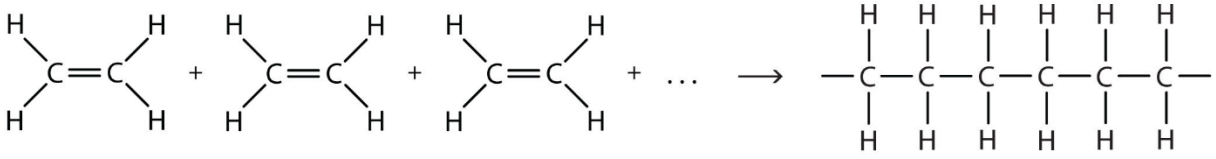
Addition polymers

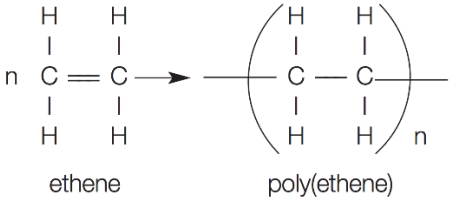
General equation



The simplest of the addition polymers is polyethene.

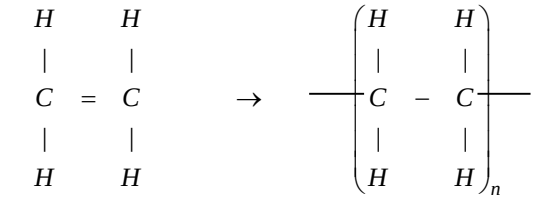
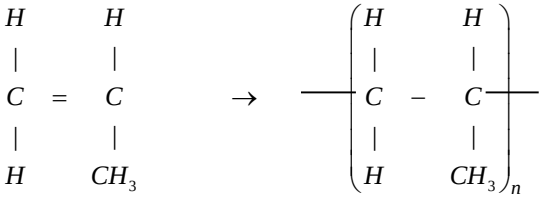
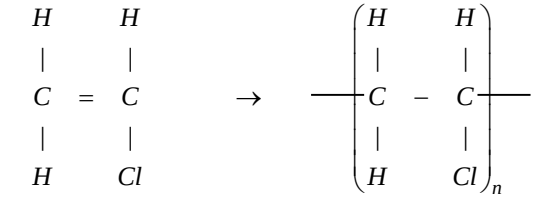
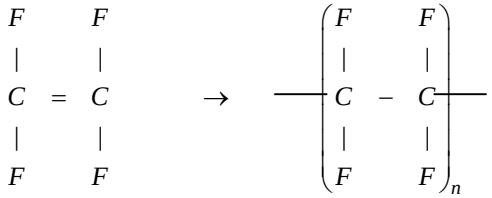
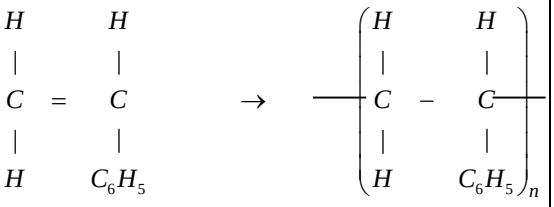
The ethene molecules are heated under high pressure in the presence of a catalyst.





CQD28.[G] Ethene, C2H4, undergoes addition polymerization.

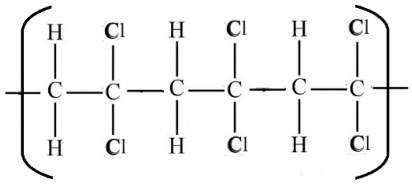
a. Draw a diagram including structural formulas to show the polymerization of ethene.
b. Identify the monomer and the repeating unit.
c. Give the name of the polymer formed.

Monomer	Name of Polymer	Equation	Use
Ethene	Poly(ethene)		Tough and durable. Easily molded into different shapes and articles such as plastic bags, bottles and buckets.
Propene	Poly(propene)		Tough and durable. Crates, boxes, rope, plastic bowls and carpet fibers
Vinyl chloride Or Chloroethene	Polyvinylchloride PVC		Stronger and harder but not as flexible as polyethene. -electrical insulation -simulated leather -waterproofing -gutters and pipes
Tetrafluoroethene	Poly(tetrafluoroethene) or PTFE or Teflon® or Tefal®		Non-stick surface and withstands high temperatures -In electrical insulation. -In making non-stick cooking utensils. -As an artificial substitute for hip joints.
Phenylethene or Styrene	Poly(phenylethene) Or Polystyrene		Light, poor thermal conductor -insulation in: battery cases and refrigerator parts. -packaging of electronic or fragile articles.

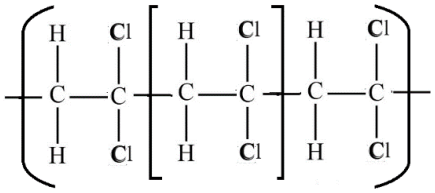
Some types of plastic can be melted down and used again. These are **thermoplastics**. Other types of plastic decompose when they are heated. These are **thermosetting plastics**. Recycling is difficult because the different types of plastic must be separated.

Deduce the monomer from the formula of the polymer

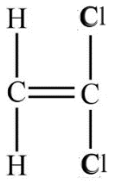
Eg. Deduce the monomer of the polymer given below,



Identify the repeating unit, it will be two carbon atoms with a single bond between them. Draw a bracket around them.



Copy the unit with a double bond between the C-atoms.

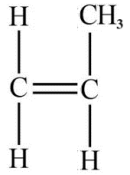


or $CH_2=CCl_2$ (1,1-dichloroethene)

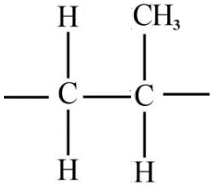
Draw the condensation polymer from its monomer:

Eg. Draw the polymer formed from propene: $CH_3CH_2=CH_2$

Draw the monomer with only the doubly bonded carbons:

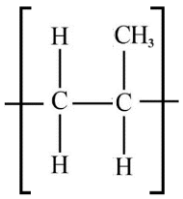


Remove the double bond and add single bonds

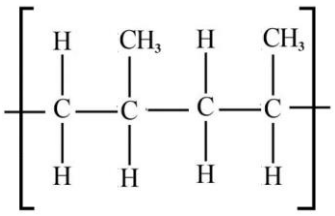


This is your repeating unit. If asked to draw

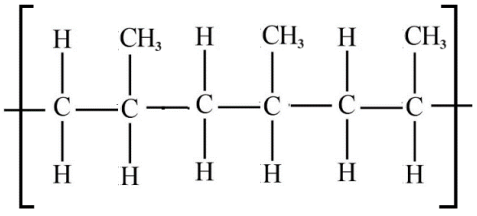
one repeating unit



two repeating units



three repeating units



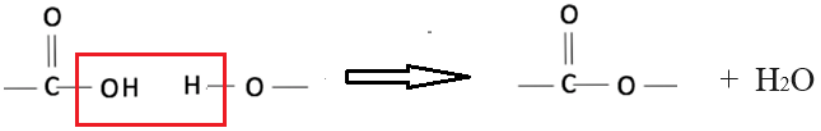
CQD23.
a. i. Write an equation to show how the following compounds polymerize.

C_2H_4; C_2F_4; C_2H_3Cl.

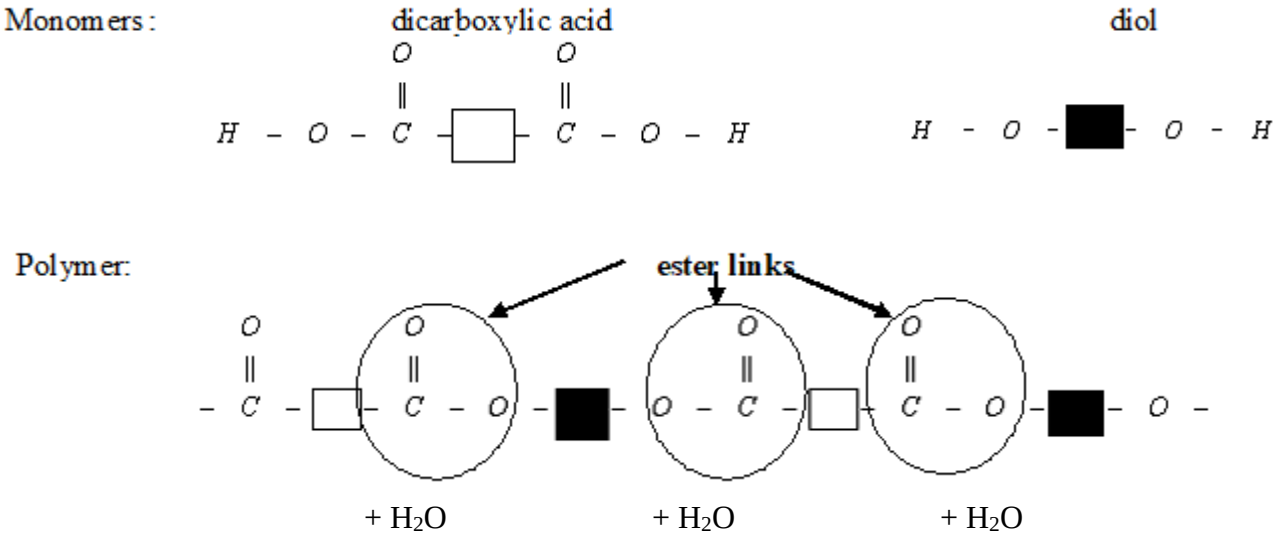
ii. Are the polymers produced addition or condensation polymers?

b. Write an equation to show how styrene, $C_6H_5C{=}CH_2$, polymerizes.

Synthetic Condensation polymers
Polyesters (acid group and alcohol group)
General equation for bond formation



Dacron or Terylene and PET (Copolymer)



NOTE: PET polymer can be converted back into monomers and re-polymerised.

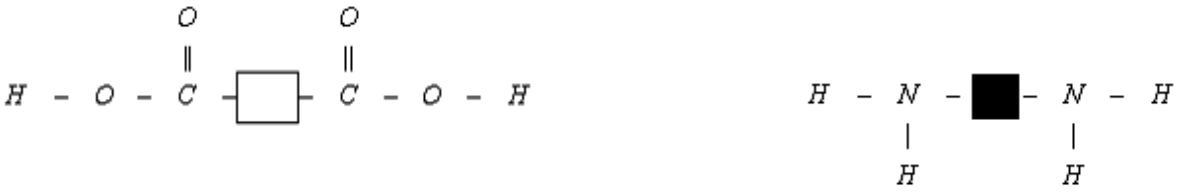
Most commonly used polymer in the world. Found in:

- many types of packaging for food and drink
- clothing fibers (textile industry)
- bottle production

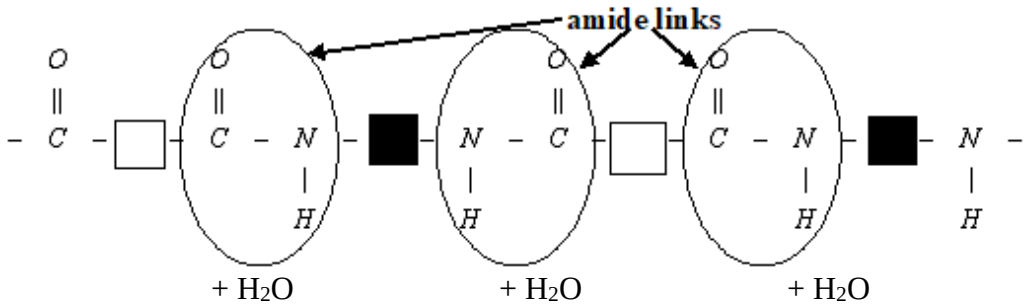
CQD29. PET is one of the most used polymers. It is used in packaging foods and beverages.
a. What are the monomers involved in the polymerization?
b. Draw the structure of a PET polymer.
c. Identify the type of the linkage formed.
d. What is a characteristic of a PET polymer?

Nylon (Copolymer)

Monomers: dicarboxylic diamine

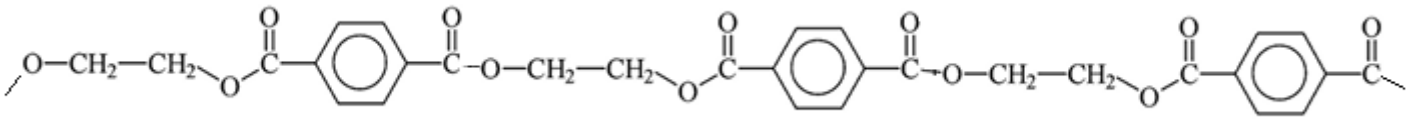


Polymer:

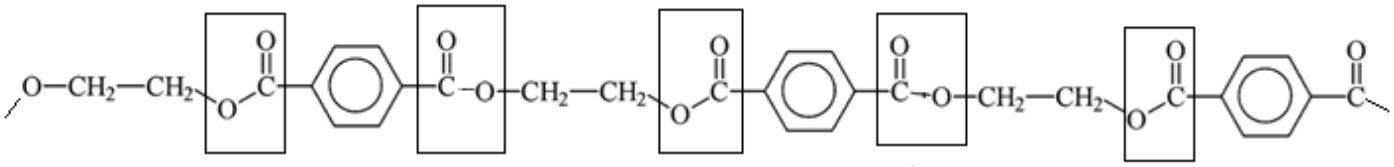


First produced as a solid, nylon is then melted and extruded through small holes to form fibers. The fibers are then stretched so the polymer molecules align and the long filaments are then dried. Once dry, the fibers are woven and used in the clothing industry, to make ropes, nets, …
Nylon can also be molded into strong durable items like gearwheels.

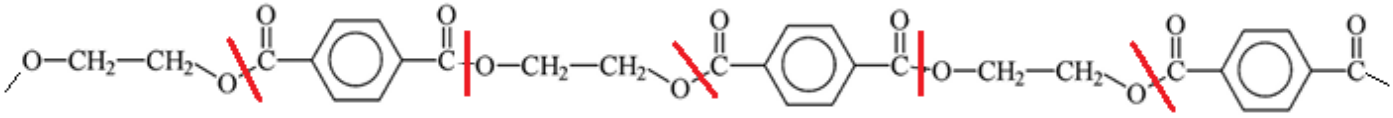
Deduce the monomer from the formula of the polymer
Eg. Deduce the monomer of the polymer given below,



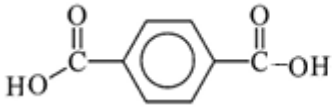
Identify the link between the monomer unit and draw lines through it.

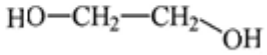


an ester bond, therefore the monomers are a carboxylic acid and an alcohol

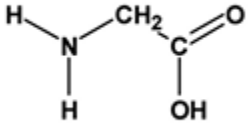


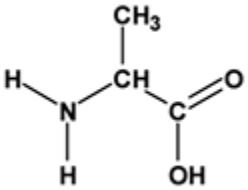
Redraw the monomers adding an H to the -O end and an OH to the C=O end.



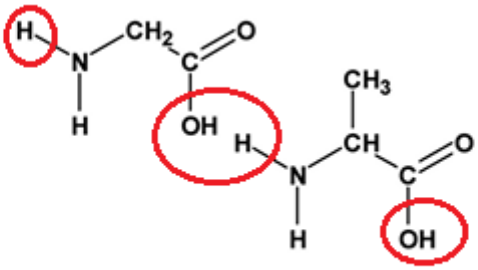


Draw the addition polymer from its monomer:
Eg. Draw the polymer formed from the following monomers:



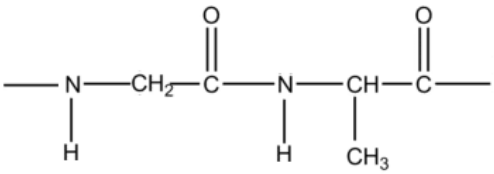


Identify the functional groups and the link they would form, then draw them with the functional groups next to each other. Circle the OH of the carboxylic group and the H of amine group:



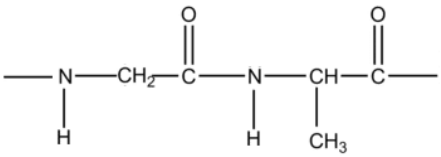
A carboxylic acid group and an amine will form an amide link.

Remove the circled groups as water and attached the two parts

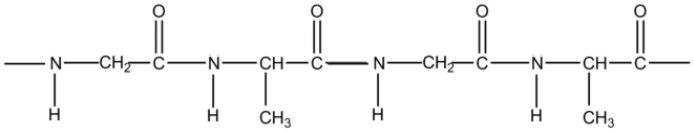


This is your repeating unit. If asked to draw

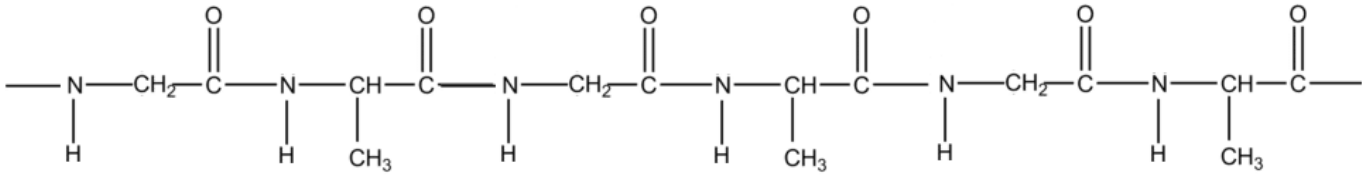
one repeating unit



two repeating units



three repeating units



Comparison between addition and condensation polymerization

	addition polymerization	condensation polymerization
Monomers	Usually, one unsaturated monomer	Usually, two monomers with at least two functional groups per molecule
Reaction type	Mononmers join together with double bond opening in an addition reaction	Every time two molecules join, a small molecule is lost in a condensation reaction
Product properties	Single product	Two products
	Non-biodegradable	Can be biodegradable (though may take a long time, for example nylon would need between 40 to 50 years to biodegrade)
	Resistant to acid/base	Can be hydrolyzed back to monomers if heated in acid/base

CQD27. [G] Give the main differences between addition and condensation polymerization.

Advantages of Plastics

1. Light (low density)
2. Extremely strong
3. Cheap
4. Resistant to corrosion.
5. Low conductivity

Disadvantages of Plastics (Problems with Plastics)

1. Plastics or synthetic polymers are very difficult to dispose of, they are usually buried in landfill sites.
2. They tend to accumulate in oceans
3. Most of plastics are **non-biodegradable** (they cannot be decomposed by bacteria in the soil).
4. Burning plastics produces toxic gases such as hydrogen cyanide HCN.
5. Difficult to recycle.

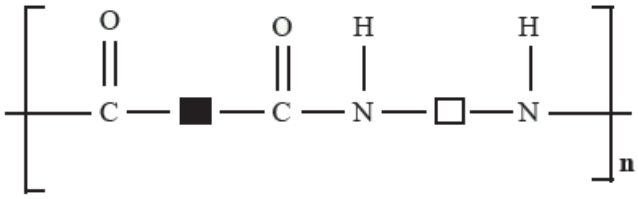
Measures to reduce impact on environment

1. Reduce plastic packaging
2. Reduce use of "single-use plastics
3. Reuse and recycle whenever possible

CQD24. Consider the following two monomers: HO-C_2H_4-OH and HOOC-C_6H_4-COOH.

a. Represent the structure of the polymer that can be produced when the monomers react together. You can use X and Y to represent certain groups of atoms.
b. Is the polymer produced an addition or a condensation polymer?
c. Circle the ester group in the polymer.

CQD25. The structure of nylon can be represented as such.

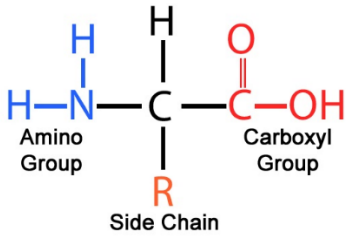


In the above structure, ■ represents a $(CH_2)_4$ and □ represents a $(CH_2)_6$ group.

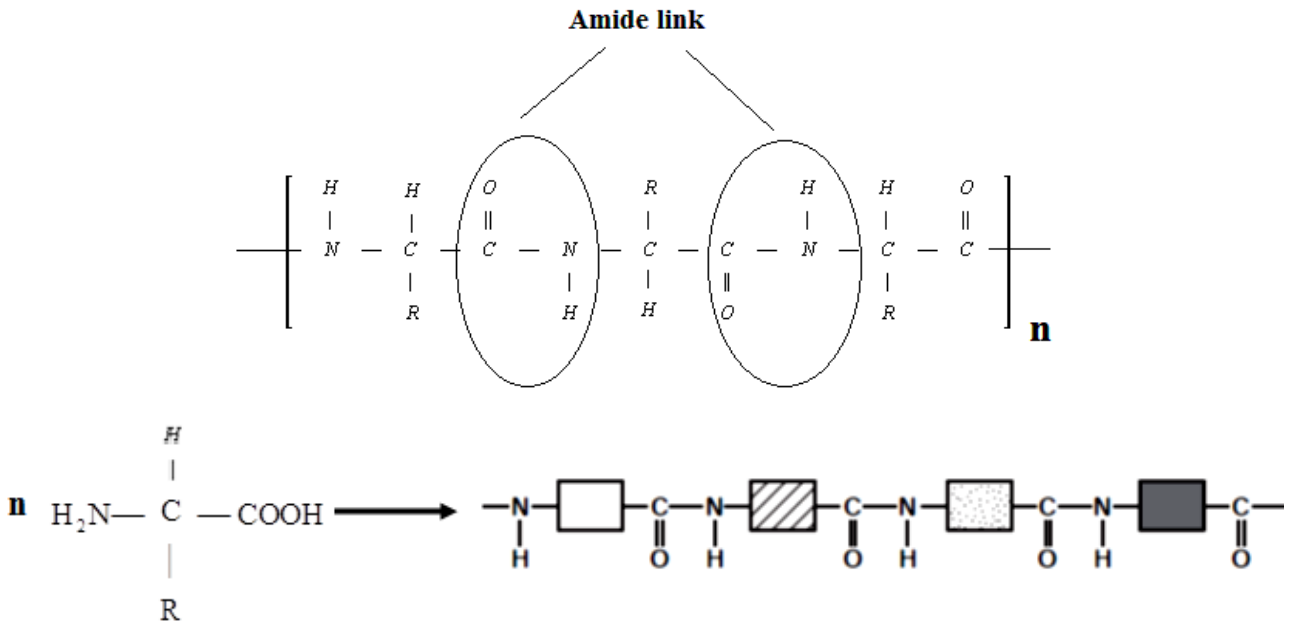
a. Is nylon an addition or condensation polymer?
b. Circle the amide bond or the peptide linkage in the above structure.
c. Draw the structures of the monomers from which nylon is produced.

<u>Natural Macromolecules:</u> Proteins (natural polyamides)

Monomer: Amino acid H_2N— R —COOH



Amino acids have a carboxylic acid group on one end and an amine group on the other.
The ends of the molecules react together to form an amide link, the same one found in nylon.



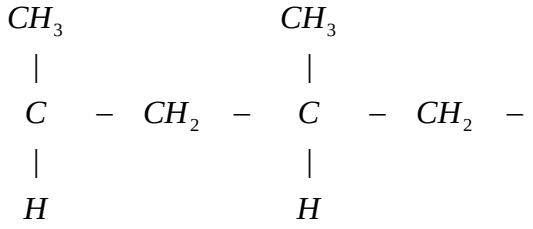
Hydrolysis of Proteins
Boil protein in acid or mix it with an enzyme: protein $\rightarrow$ amino acids

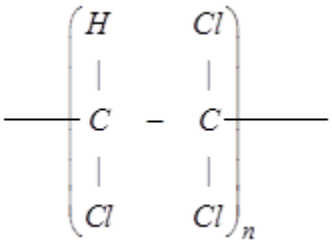
How can you separate and identify amino acids present in a protein?
Protein is hydrolyzed, in the presence of dilute hydrochloric acid, into amino acids.
Amino acids are separated by chromatography.

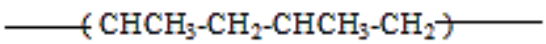
Identification of the amino acids is done by:

1. **Spraying the chromatogram with a suitable locating agent as resorcinol or ninhydrin.**
2. **Measuring the R_F value and compare the results to the data booklet.**

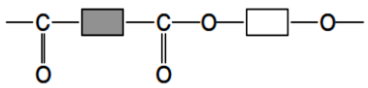
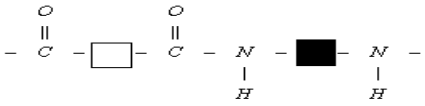
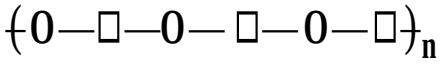
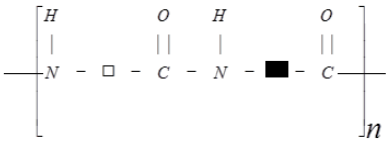
1. Give the monomer of the following polymers:

a.  **Answer:**

b.  **Answer:**

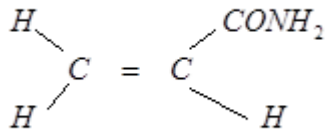
c.  **Answer:**

2. Specify the following polymers and their corresponding monomers:

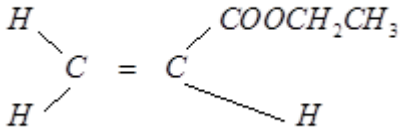
Structure of polymer	Type of polymer	Corresponding monomer
		
		
		
		

3. Draw the polymer formed by the following monomers and specify whether it is an addition or a condensation polymer:

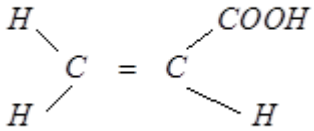
   1. Acrylamide



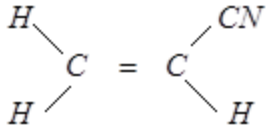
   2. Ethylacrylate



   3. Propenoic acid



   4. Acrylonitril



Topic 6-Nitrogen

Nitrogen is obtained industrially by the fractional distillation of liquid air. It is a highly unreactive gas and can only be used by plants once nitrogen fixing bacteria change it into soluble compounds in the soil.

CQD 1.
a. How is nitrogen extracted industrially?
b. Does the process involve primarily physical or chemical changes?
c. Is the nitrogen collected pure?

Uses of nitrogen

1. In the industrial manufacture of ammonia by the Haber Process
2. Making fertilisers

CQD 2. List two uses of nitrogen.

Haber Process: Industrial manufacture of ammonia gas
Ammonia is a colorless gas with a distinctive odor. It is less dense than air and very soluble. It is very important in making fertilizers and explosives. This is why its production is a very important industrial process.

Ammonia is produced by the direct reaction between nitrogen and hydrogen gases in a reversible reaction.

Raw materials

- Nitrogen from the fractional distillation of liquefied air
- Hydrogen from: Methane, $CH_4(g) + 2H_2O(g) \longrightarrow CO_2(g) + \mathbf{4H_2(g)}$

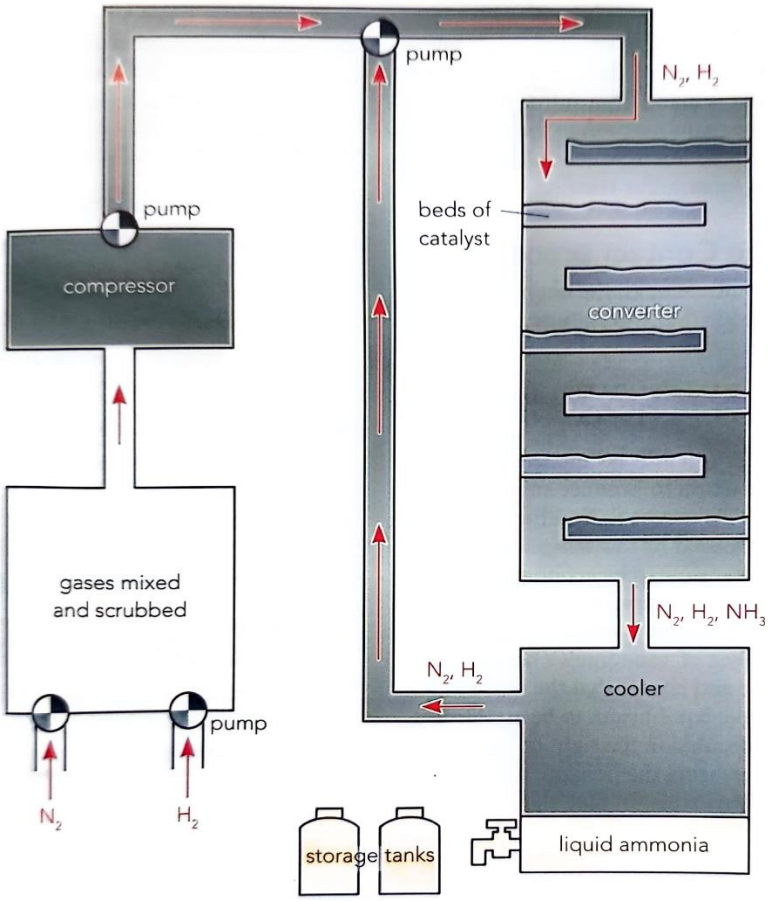
Or $CH_4(g) + H_2O(g) \xrightarrow{Ni\ catalyst} CO(g) + \mathbf{3H_2(g)}$
Followed by reacting CO(g) further with more steam: $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2O(g)$

Conditions of the Haber Process

Temperature: 450°C
Pressure: 200 atm (**20,000kPa**)
Catalyst: finely divided iron, Fe

$$N_2(g) + 3H_2(g) \underset{450℃\ and\ 200\ atm}{\overset{finely\ divided\ iron}{\rightleftharpoons}} 2NH_3(g) \qquad \Delta H = -92\ kJ/\ mole$$

Being such an important process, the Haber process has been studied over a very wide range on temperature and pressure conditions.



From Chemistry for Cambridge IGCSE Coursebook

<u>The effect of altering the conditions on the rate of the reaction and the reaction yield</u>

1. Temperature

<u>Increasing temperature</u> favours the endothermic reaction so equilibrium shifts to the left.
Result: Lower percentage yield and faster rate of reaction.

2. Pressure

<u>Increasing pressure</u> favours the lesser # of gaseous moles so equilibrium shifts to the right.
Result: Higher percentage yield and higher reaction rate
Problem: High pressure is an expensive procedure and there might be a high risk of explosion.

- **<u>Important Notes</u>**
  - During the Haber process the percentage yield of ammonia is low, 8%.
  - Although the Haber conditions of temperature (450℃) and pressure (20,000 kPa) give a low percentage yield of ammonia, we still use them as a compromise or optimum conditions to produce enough product at an acceptable rate. These conditions are more economic than getting a higher percentage yield of ammonia at a more costly high pressure and a slower lower temperature reaction.

3. Catalyst: Faster reaction rate but no effect on yield.

4. Reducing the concentration of ammonia:
Ammonia has a much higher boiling point than nitrogen and hydrogen gases, so it can be easily removed from the reaction mixture by cooling the mixture when the % of ammonia in it reaches around 15%. The cooling will condense the ammonia into a liquid, leaving hydrogen and nitrogen gases which get recycled. This will pull the equilibrium system toward producing more ammonia to replace the removed ammonia. By recycling the yield can be increased to about 98%.

Ammonia is stored under pressure as a liquid.

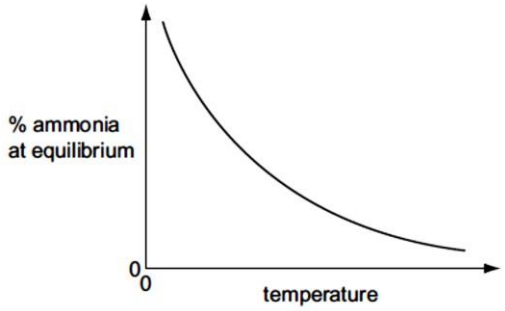
Ammonia can be produced in the laboratory when ammonium salts react with bases. Typical example would be heating ammonium chloride and calcium hydroxide:

$$2NH_4Cl(s) + Ca(OH)_2(s) \rightarrow 2NH_3(g) + CaCl_2(s) + 2H_2O(l)$$

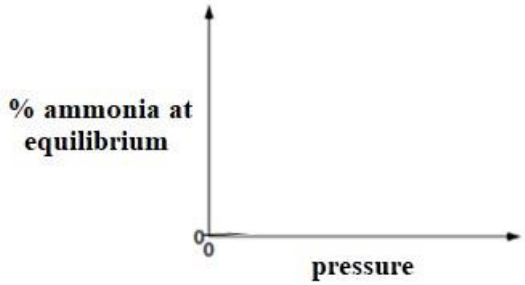
CQD 4. Ammonia, NH_3, is produced by the Haber process according to the following chemical equation.

$$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$$

The graph below shows how the percentage of ammonia changes with temperature in the mixture at equilibrium.



a. What is meant by the term 'dynamic equilibrium'?
b. How does the percentage of ammonia change with temperature?
c.
i. On the graph below, sketch the graph that shows what happens to the percentage of ammonia in the equilibrium mixture as the pressure changes.



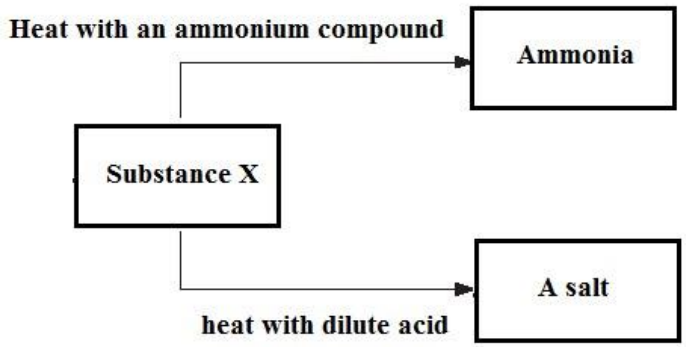
ii. Explain why the graph in part c has the shape described.

CQD 5. Answer the following questions regarding the industrial manufacture of ammonia - the Haber Process.

a) During the Haber process, only a small percentage of hydrogen and nitrogen gas is converted into ammonia. How are the unreacted gases separated from ammonia?
b) What happens to the unreacted nitrogen and hydrogen after it has been separated from ammonia?
c) Haber process uses 450°C resulting in a percentage yield of 15%. Explain.
d) What is the effect when the temperature is increased in the Haber process,
   i. on rate of reaction
   ii. on position of equilibrium?
e) What is the effect on equilibrium position when the pressure is increased in the Haber process? Explain your answer

CQD 6.

a. The diagram shows some reactions of substance X.



Which type of substance is X?

b. How could we obtain ammonia from ammonium sulfate?

A. Heat it with an acid.
B. Heat it with an alkali.
C. Heat it with an oxidizing agent.
D. Heat it with a reducing agent.

c. Complete the following equation: … …+ $NH_4Cl \rightarrow NaCl + NH_3 + H_2O$

CQD 15. [G] Ammonia, NH_3, is produced industrially from nitrogen gas, N_2, and hydrogen gas, H_2, by a process called the Haber process.

a. Give the chemical equation for the synthesis of ammonia in the Haber process.
b. What are the conditions of pressure and temperature needed for optimum yield of ammonia in this process?
c. Give the name of catalyst used in this process.

Fertilizers

Plants usually obtain what they need to make their food by photosynthesis from the soil and air. However, when the same plot of land is used every year, nutrients in the soil get depleted and need to be replenished by fertilizers. Fertilizers need to be water soluble to be accessible to plant as they absorb them through their roots.

Ammonia is used in the production of fertilizers.
Ammonium salt and nitrates are used as artificial fertilizers. The most important of which is ammonium nitrate which is the principle nitrogenous fertilizer containing 35% by mass nitrogen.

$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$$

The ammonium nitrate product is dried and spread on the land.

Fertilizers usually contain compounds of the three essential elements: nitrogen *(protein production and green foliage)*, potassium *(healthy growth of flowers and fruits)* and phosphorus *(strong roots*). Such fertilizers are commonly known as "NPK" fertilizers.

1. **Nitrogen (N)**
   Examples: ammonium nitrate, NH_4NO_3 or ammonium phosphate, $(NH_4)_3PO_4$
2. **Phosphorus (P)**
   Examples: ammonium phosphate, $(NH_4)_3PO_4$
3. **Potassium (K)**
   Examples: potassium phosphate K_3PO_4 or potassium nitrate, KNO_3.

Different plants require different proportions of these elements.

Importance of fertilizers

1. Increase rate of plant growth
2. Higher yield for crops to feed the world's growing population

Problems with fertilizers

If too much fertilizer is applied to the land, rain washes the fertilizer off the land and into rivers and streams. This leaching leads to the growth of algae and marine plants. As the algae die and decay, oxygen is removed from the water, leaving insufficient amounts for fish and other organisms to survive.

CQD 3.

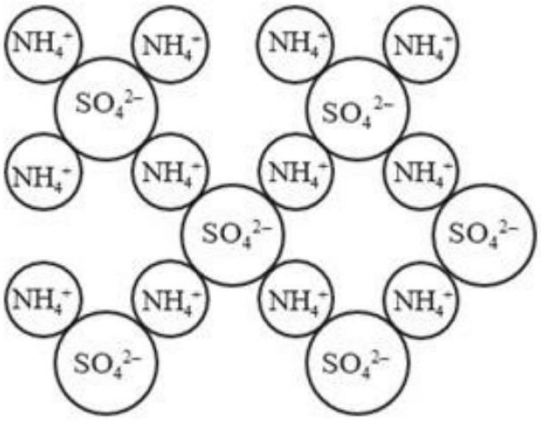
a. Give the most important use of ammonia?
b. Ammonia is used to make ammonium nitrate, NH_4NO_3, by reacting ammonia with nitric acid. Write a chemical equation for the reaction between ammonia and nitric acid.

CQD 13.

b. What is an NPK fertilizer? Why do plants need fertilizers?
c. List one advantage and one disadvantage of using fertilizers.

CQD 8. Fertilizers usually contain compounds of nitrogen, phosphorus and potassium - nitrogen for *green foliage*, potassium for *healthy growth* and phosphorus *for strong roots.*

a. Ammonium phosphate, $(NH_4)_3PO_4$, is a fertilizer. Which **two** elements in ammonium phosphate are *important* for plant growth?
b. Why do farmers use fertilizers?
c. Many fertilizers contain ammonium sulfate. Ammonium sulfate is made by reacting aqueous *ammonia* with sulfuric acid. What is the type of this reaction?
d. Aqueous ammonia reacts with nitric acid to make another compound often found in fertilizers. *Give* the name of the compound produced.
e. The structure of ammonium sulfate is shown below.



Give the simplest ratio of ammonium and sulfate ions in ammonium sulfate.

f. Ammonium salts react with alkalis. For example:
ammonium sulfate + sodium hydroxide → sodium sulfate + ammonia + water

Deduce why adding slaked lime to fields which have fertilizers spread on them may result in loss of nitrogen.

CQD17. [G] When petrol burns in a car engine, pollutant gases are produced. Name two gas pollutants that are released by a combustion engine and are converted into less harmful products in catalytic converters installed in a car exhaust.

CQD 19. [T] Which of the following fertilizers have the highest percentage by mass of sulfur?

A. K_2SO_4
B. $CaSO_4$
C. $Al_2(SO_4)_3$
D. $(NH_4)_2SO_4$
E. $MgSO_4$

Topic 8-Sulphur

Contact Process: Industrial manufacture of sulphuric acid

Sulfur dioxide is obtained by burning sulfur or roasting sulfide ores (from the volcanic deposits in Iran, Iraq, Poland and Mexico)
Oxygen is obtained from the fractional distillation of liquefied air.

Sulphur → sulphur dioxide → sulphur trioxide → sulphuric acid

In a first step, sulfur is burnt (or sulfide ores, such as ainc blende, are roasted) in oxygen:

$$S(s) + O_2(g) \xrightarrow{heat} SO_2(g)$$

$$\text{Metal sulfide + oxygen} \xrightarrow{heat} \text{metal oxide + sulfur dioxide}$$

$$2ZnS(s) + 3O_2(g) \xrightarrow{heat} 2ZnO(s) + 2SO_2(g)$$

Raw materials:

- Sulfur: mined from various parts of the world. Today a major source is from fossil fuel impurities during desulfurization.
- Oxygen: from fractional distillation of air

In a second step, sulfur dioxide is cleaned and mixed with oxygen, then passed into a tower containing vanadium (V) oxide pellets in trays. The gases react reversibly to form sulfur trioxide in a 98% yield.

$$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$$

Raw materials:

- Sulfur dioxide form burning sulfur or roasting sulfides.
- Oxygen: from fractional distillation of air

Conditions of the Reaction

- Temperature: 450°C
- Pressure: 2 atm (**200 kPa**)
- Catalyst: Vanadium (V) oxide, V_2O_5

Sulfur trioxide produced is then dissolved into 98% concentrated sulphuric acid to form oleum (fuming sulphuric acid). The SO_3 is NOT dissolved in water becaue the reaction is highly exothermic and to acid mist.

$$SO_3(g) + H_2SO_4(l) \rightarrow H_2S_2O_7(l)$$

Oleum is transported and then added to water to produce sulphuric acid of the required concentration on site. (Oleum is always added to water and not the other way around because the reaction is highly exothermic and the acid could spit out of the container causing severe burns).

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$

Unreacted gases are recycled.

The effect of altering the conditions on the rate of the reaction below and the reaction yield

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{Energy}$$

1. Temperature

Decreasing temperature favours the exothermic reaction so the equilibrium shifts to the right.
Result: Higher percentage yield and slower rate of reaction.

2. Pressure

Increasing pressure favours the lesser number of gaseous moles so the equilibrium shifts to the right.
Result: Higher percentage yield and higher reaction rate
Problem: High pressure is an expensive procedure and there might be a high risk of explosion.

3. Catalyst

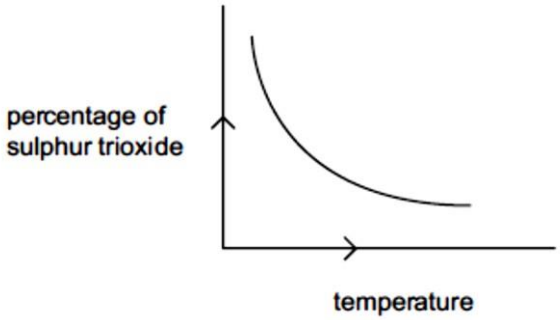
Result: Faster reaction rate but no effect on yield.

CQD 1.
a. List the raw materials that are used in the industrial manufacture of H_2SO_4.
b. List the conditions used in the Contact process.
c. Write an equation for the catalysed reaction.

CQD 2. The equation for the second step in the Contact process is:

$$2SO_2 + O_2 \rightleftarrows 2\ SO_3$$

The percentage of sulfur trioxide in the equilibrium mixture changes with temperature.



a. i. How does the percentage of sulfur trioxide change in the equilibrium mixture as the temperature increases?
ii. What is the effect of increasing temperature and increasing pressure on the position of equilibrium?

b. In the following reaction $2SO_2 + O_2 \rightleftarrows 2SO_3$, is the forward reaction exothermic or endothermic? Explain your choice.

c. The temperature used in the Contact process is 450°C. Explain in terms of rate and percentage yield why is this temperature used.

d. i. Explain how sulfur trioxide is converted to concentrated sulfuric acid.
ii. Complete the following equations regarding the preparation of sulfuric acid from sulfur trioxide:

$$SO_3 + \_\_\_\_\_ \longrightarrow H_2S_2O_7$$

$$H_2S_2O_7 + \_\_\_\_\_ \longrightarrow H_2SO_4$$

CQD 3.
a. Consider the following equations that show a possible route for the formation of sulfuric acid as part of acid rain.

$$SO_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow SO_3(g)$$

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

Which of the two equations above represents a redox reaction?

b. An alternative way to produce acid rain is through the reaction of SO_2 with water (rain) first, then through the reaction of the product with oxygen gas. Write the two equations that are involved in this process

CQD 4. [G] Sulfuric acid, H_2SO_4, is manufactured industrially by the Contact process. After reacting molten sulfur with air in a large furnace, the sulfur dioxide gas, SO_2, produced is mixed with more air to make sulfur trioxide gas, SO_3.

a. Give the chemical equation for the reaction between sulfur dioxide and air to form sulfur trioxide.

b. What are the conditions of pressure and temperature to produce SO_3 in this process?

c. Give the name of the catalyst used in this reaction.

CQD 5. [T] The manufacture of sulfuric acid by the Contact Process is done by the following reactions.

Reaction 1: Sulfur is changed to sulfur dioxide gas.

Reaction 2: Sulfur dioxide combines with oxygen and from sulfur trioxide gas. This step is reversible and a catalyst is used to speed it up.

Reaction 3: Sulfur trioxide is dissolved in concentrated sulphuric acid solution. Oleum, fuming sulphuric acid ($H_2S_2O_7$), forms.

Reaction 4: Oleum is then diluted with water to form concentrated sulfuric acid.

a. Define a catalyst. Name the catalyst used in reaction 2.
b. Write a balanced equation for reaction 2.
c. Write a balanced equation for reaction 4.
d. In reaction 2, less sulfur trioxide forms at a higher temperature. Specify whether reaction 2 is endothermic or exothermic. Justify your answer.

Topic 9-Halogens

They are group VII elements located on the right hand side of the Periodic Table like other non-metals.

Halogen	Symbol
Fluorine	F
Chlorine	Cl
Bromine	Br
Iodine	I
Astatine	At

General Physical Properties of Halogens

1. Group VII elements, halogens, are diatomic non-metals. Below are some properties at RTP.

Halogen	Molecule	Physical state	Color	
Fluorine	F_2	Gas	Pale yellow-green	
Chlorine	Cl_2	Gas	Yellowish green	
Bromine	Br_2	Liquid	Red- brown	brown gas
Iodine	I_2	Solid	Grey-black	purple gas
Astatine	At_2	Solid (Radioactive)	Black	

2. Poor conductors of heat and electricity (because they are non-metals, have no free ions or electrons that can move and carry electric current).
3. They are of low solubility in water.
4. All are toxic
5. They show they following trends down the group:
   - Increase in density
   - Decrease in reactivity

- **Test for Chlorine gas**

  Test: Approach a damp litmus paper:
  Observations: If the paper is blue litmus, it turns red then bleaches.
  If the paper is red litmus, it bleaches.

CQD1.

a. What is the color of chlorine gas? Is it lighter or denser than air?
b. How do you test for chlorine gas?

- **Test for halides**

Add few drops of silver nitrate, $AgNO_3$, acidified with nitric acid to the aqueous solution

Chloride	Bromide	Iodide
Observation A white precipitate of silver chloride, AgCl, is formed, which darkens upon exposure to light. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$	Observation A creamy precipitate of silver bromide, AgBr, is formed, which darkens upon exposure to light. $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$	Observation A yellow precipitate of silver iodide, AgI is formed, which darkens upon exposure to light. $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

- **Reactivity of Halogens**

Fluorine is the most reactive halogen. It is the most reactive non-metal.
The reactivity of the halogens decreases as we move down the group.

1) Displacement reactions of Halogens

Generally, $X_2(aq) + 2Y^-(aq) \rightarrow 2X^-(aq) + Y_2(aq)$

If X is Cl, then Y is Br or I
If X is Br, then Y is I
The halogen that is displaced dissolves in water and colors the solution

Possible reaction	Observation
$Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$	The solution changes from colorless to orange due to the formation of aqueous bromine.
$Cl_2(g) + 2NaI(aq) \rightarrow 2NaCl(aq) + I_2(aq)$	The solution changes from colorless to dark reddish brown due to the formation of aqueous iodine.
$Br_2(l) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(aq)$	The solution changes from colorless to dark reddish brown due to the formation of aqueous iodine.

CQD6. Write the net ionic equation for the reaction that takes place when a few drops of silver nitrate solution are added separately to the aqueous solutions listed below.
State the observations you expect to make in each case.

a. Sodium bromide
b. Potassium iodide
c. Sodium chloride
d. Potassium fluoride

CQD7.

a. Complete and balance the following ionic equations.
Where no reaction is expected, write NR.

i. $Cl_2(g) + Br\text{-}(aq) \rightarrow$
ii. $I_2(aq) + Br\text{-}(aq) \rightarrow$
iii. $F\text{-}(aq) + Br_2(aq) \rightarrow$
iv. $I\text{-}(aq) + Cl_2(aq) \rightarrow$

b. What are the colors of bromine and iodine solutions in cyclohexane? (an organic solvent similar to CCl_4)

CQD8. [G] Use the trend in reactivity of halogens to predict the spontaneity of each of the following reactions under normal conditions.

a. $Cl_2 + 2KI \rightarrow I_2 + 2KCl$
b. $I_2 + 2KF \rightarrow F_2 + 2KI$
c. $Br_2 + 2KCl \rightarrow Cl_2 + 2KBr$

CQD12. [T] Which sequence best describes the relative reactivity of halogens from most reactive to least reactive?

A. chlorine, bromine, iodine
B. chlorine, iodine, bromine
C. bromine, chlorine, iodine
D. bromine, iodine, chlorine
E. iodine, bromine, chlorine

Topic 10-The Metals

General Physical Properties of metals

1. Good conductors of electricity
2. Good conductors of heat
3. Solids (except mercury) at RTP with fairly high melting point.
4. Most have high density (except Alkali metals)
5. Ductile (can be drawn into a wire)
6. Malleable (can be hammered into shapes)
7. Silvery luster (except copper and gold)

Metals Versus Non-Metals

Property	Metals	Non-Metals
Melting and boiling points	Solids (except mercury) at RTP with fairly high melting and boiling point	Solids or gases (except bromine) at RTP with usually low melting and boiling point
Electrical conductivity	Good	Poor (except graphite)
Thermal conductivity	Good	Poor (except diamond)
Malleability and ductility	both	Neither
Strength and hardness	Usually strong and hard with generally high densities	Most are soft (except diamond) with low densities
Ability to produce sound	Sonorous (rings like a bell when hit with a hammer).	Not sonorous
Appearance	Grey (except copper and gold) Can be polished to give a shiny surface	Have different colors, with usually dull surfaces

Group I alkali metals

They are relatively soft metals, easily cut by a knife, that exhibit general trends down the group:

- Reactivity increases (exemplified by their reaction with water)
- Melting points decrease
- Density increases

They have relatively low densities and melting points compared to other metals.
They have shiny lustrous surface when freshly cut, but quickly tarnish when exposed to air.

CQD1.
a. List four physical properties of alkali metals.
b. Why are alkali metals never found in their pure form in nature?

CQD2. You are supplied with three samples of white powders and told that they are the nitrates of lithium, sodium, and potassium. Explain how you can identify each powder.

Transition Metals

(a) Physical properties

Are harder, denser and have higher melting and boiling points than other metals.

(b) Chemical properties:

1. **They form colored compounds**

- $CuSO_4.5H_2O$ blue
- $Fe(OH)_3$ rusty brown
- $K_2Cr_2O_7$ orange
- $KMnO_4$ purple

2. **Form ions with multiple oxidation states**

- Copper Cu^{1+} and Cu^{2+}
- Iron Fe^{2+} and Fe^{3+}
- Cobalt Co^{2+} and Co^{3+}
- Manganese Mn^{2+} and Mn^{4+}

3. **Have catalytic properties (As elements and as compounds).**

- Iron (in the finely divided form) used in the Haber process for manufacture of ammonia.
- Vanadium in the form of V_2O_5 is used in the Contact process for the manufacture of sulphuric acid.
- Manganese (IV) oxide MnO_2 is used in the preparation of oxygen from hydrogen peroxide
- Nickel is used in the hydrogenation of alkenes to produce alkanes.

CQD10. What are transition metals?

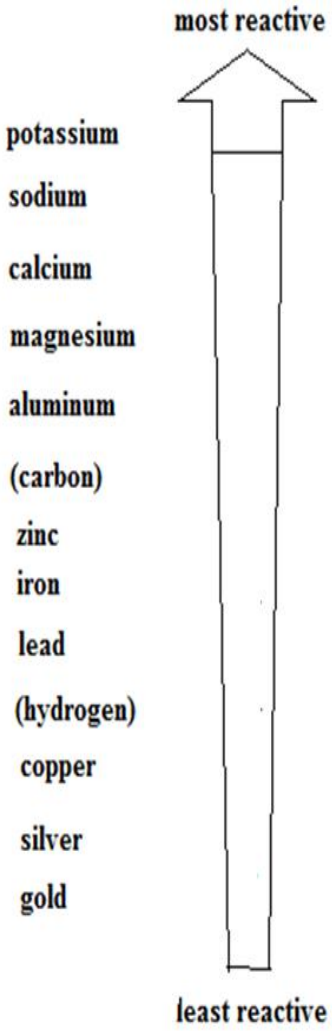
CQD29. [G] List six characteristics of transition metals.

Reactivity series:

Transition metals are less reactive than Group I and II

Metals are arranged in a decreasing order of their reactivity (in decreasing order of the tendency of a metal to form its positive ion.)

The more reactive a metal is, the easier it forms its positive ion and the harder it is to thermally decompose its compounds.



Displacement reactions of metals and their compounds can be used to work out the reactivity series. For example:

Experiment I

zinc + copper (II) sulphate → zinc sulphate + copper

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

Observations:

1. Blue color of the solution fades gradually till it disappears.
2. Red brown solid is deposited (copper metal).

Since zinc can replace Cu^{2+} from its solution, then zinc is more reactive than copper.

Experiment II

copper + silver nitrate → copper (II) nitrate + silver

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Observations:

1. Silvery deposit appears (silver metal)
2. The colorless solution turns blue due to the formation of Cu^{2+}ions.

Since copper can displace Ag^+ ions from its solution, then copper is more reactive than silver.

From the results of the two experiments above we can place the three metals in the following order of reactivity: Zn, Cu, Ag (Zn is being the most reactive).

After many such experiments the reactivity series was built.

Reactions of Metals with Oxygen

Some metals react immediately once exposed to air forming oxides, like alkali metals.
Some need to be powdered and heated, like iron. However, in the presence of water, iron will slowly react with oxygen in air forming the oxide we call rust (hydrated iron(II) oxide). Rust does not stick to the iron surface but flakes off. Aluminum reacts also very quickly with oxygen forming its oxide. However, unlike iron(III) oxide, it sticks to the metal surface "protecting it from further reacting. This makes aluminum "appear" less reactive than it actually is.

The reaction with oxygen can be used to compare reactivity of metals. For example, Li, Na and K all react with air forming an oxide, however, Li tarnishes slowly, Na quickly and K very quickly.

Metal	Reaction with oxygen
Potassium	Burns brightly in air with a purple (or lilac) flame **$4K(s) + O_2(g) \rightarrow 2K_2O(s)$**
Sodium	Burns brightly in air with a yellow flame **$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$**
Calcium	Burns brightly in air with a brick red flame **$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$**
Magnesium	Burns in air with a bright white flame **$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$**
Aluminum	Burns less strongly provided the oxide layer is removed **$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$**
Zinc	Burns less strongly **$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$**
Iron	Burns less strongly **$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$**
Lead	Reacts slowly if heated **$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$**
Copper	Does not burn, but slowly forms a black oxide layer **$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$** Reddish-brown black
Silver	No reaction
Gold	

CQD6. [G] Write the chemical equation for each of the following reactions.
a. Sodium metal reacting with oxygen gas.
b. Magnesium metal reacting with oxygen gas.
c. Aluminum metal reacting with oxygen gas

Reactions of Metals with Water /Steam

- Metals that react with water cold $H_2O(l)$ to give hydrogen gas: K, Na, Ca and Mg
  In general: ***M + $H_2O(l)$ or $H_2O(g)$ → MOH (aq) + $H_2(g)$***
  Metal water or steam metal hydroxide hydrogen

  NB. Mg reacts very slowly with cold water and faster with steam

- Metals that react with steam: Al, Zn, Mg and Fe
  In general ***M + $H_2O(g)$ → MO(s) + $H_2(g)$***

The reaction with water can also be used to compare reactivity of metals. For example Li, Na and K all react with water producing the hydroxide and hydrogen gas, however Li reacts steadily, Na strongly and K violently.

Generally, reactivity of metals decreases down a group and Group II metals are less reactive than Group I.

CQD7. List some of the chemical properties of the alkaline earth metals. Say whether they react with cold water, and how their reactivity varies down the group.

CQD8. You are supplied with a sample of white powder and told that it is calcium chloride. Explain how you can identify the cation in the powder.

CQD9.

a. Write equations for the reactions of calcium with:

i. water.

ii. oxygen.

b. Write an equation to show how limewater is used to test for carbon dioxide gas. State any expected observation.

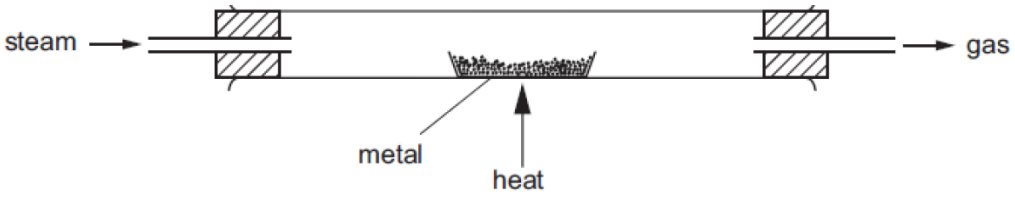
Reactions of Metals with Acid

Moderately reactive metals react safely with acids forming the metal salt and releasing hydrogen gas.

In general: ***Metal + acid → metal salt + hydrogen***

Metal	Reaction with water	Reaction with dilute acid
Potassium	Violent reaction in cold water, forming hydrogen which then burns; with a lilac flame (due to potassium ions) **$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$**	Very violent and dangerous reaction, which should never be tried in the laboratory **$2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$**
Sodium	Violent reaction in cold water, forming hydrogen **$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$**	Very violent and dangerous reaction, which should never be tried in the laboratory **$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$**
Calcium	Reacts well with cold water forming a white precipitate of calcium hydroxide **$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$**	Very vigorous reaction **$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$**
Magnesium	Very slow reaction with cold water but burns readily in steam **$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$**	Vigorous reaction, gas bubbles, Mg disappears and colorless solution forms **$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$**
Aluminum	Reacts with steam when heated (provided oxide layer removed) **$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$**	Slow to react with cold acid, bubbles form when heated, Al disappears and colorless solution forms **$4Al(s) + 6HCl(aq) \rightarrow 4AlCl_3(aq) + 3H_2(g)$**
Zinc	Reacts when steam is passed over hot metal. Zinc oxide is produced that is yellow when hot and white upon cooling. **$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$**	Bubbles form, Zn disappears and colorless solution forms **$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$**
Iron	Reacts when steam is passed over hot metal. **$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$**	Bubbles form, Fe disappears and pale-green solution forms **$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$**
Lead	No reaction	
Copper		
Silver		
Gold		

CQD11.
a. Which metals react with water, and which react with steam only? Which metals do not react with water or steam?
b. The diagram below shows the apparatus used to test the reactivity of sodium, iron and silver with steam. Which of the metals can react with steam to forms hydrogen gas?



	Sodium	iron	silver
[-A-]	√	√	√
[-B-]	x	√	√
[-C-]	√	x	√
[-D-]	√	x	x
[-E-]	√	√	x

CQD12.
a. Which metals react with dilute HCl (aq)?
b. Which are not able to react with 1.0 M HCl (aq)?

CQD13. The table below gives information about three different metals.

Metal	metal oxide is reduced upon heating with carbon	metal can react with dilute nitric acid
P	√	√
Q	√	x
R	x	√

What is the order of reactivity of the above three metals?

	least reactive ⟶ most reactive		
[-A-]	Q	P	R
[-B-]	Q	R	P
[-C-]	R	P	Q
[-D-]	P	Q	R
[-E-]	P	R	Q

CQD37. Write equations for the reaction of aluminum with:

a. air.
b. HCl(aq).
c. steam.

Observation for reaction of different metals with water

i. Reaction of Na or K or Li with water
   1. Piece of metal melts into a ball, floats on water surface and darts about
   2. A small explosion occurs with K producing a colored vapor K produces a lilac flame due to K^+ ions
   3. Piece of metal decreases in size (as it dissolves in water and pH of water changes from 7 to 14 due to formation of hydroxide ions, OH^-)

ii. Reaction of Ca with water
   1. Piece of metal sinks
   2. Metal moves up and down due to the bubbles of hydrogen gas
   3. Solution turns milky or white (due to the formation of $Ca(OH)_2$ that has a pH of 11)

Extraction of metals:

Very few metals are found in the pure state, most exist as ores containing minerals mixed with unwanted rock.

The ease of obtaining the metal from its ore is related to its position in the reactivity series. The higher up (more reactive) the more difficult to obtain the metal from its ore.

The method of extraction depends on the reactivity of the metal.

Metal	Extraction method
reactive metals	electrolysis
potassium sodium calcium magnesium	of molten chloride
aluminium	electrolysis of molten oxide
(carbon)	
zinc iron tin lead copper	Reduction of oxides: These metals are below carbon in the reactivity series and so can be obtained by heating their oxides with carbon
Silver and Gold	The least reactive metals are found as pure elements

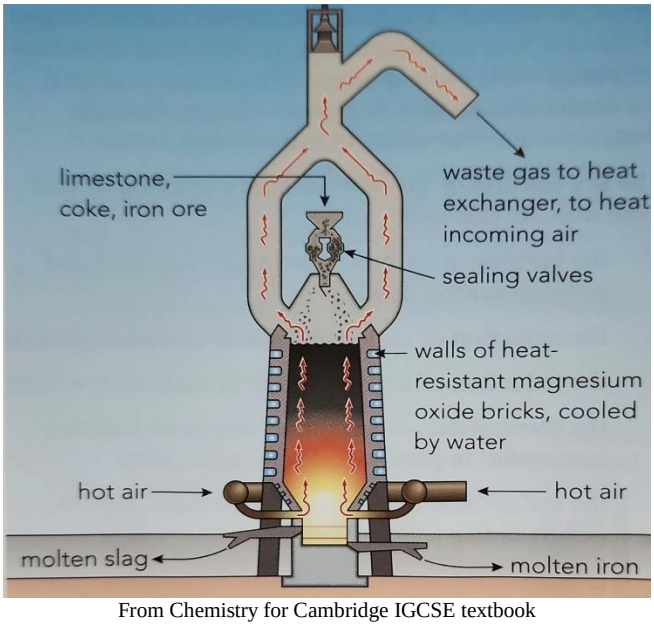
Extraction of Iron: The Blast Furnace

Iron is the second most common metal in Earth's crust (5%) after aluminum. The main ore hematite is around 60% iron.

Ore: Hematite (Iron (III) oxide, Fe_2O_3)

Method: Reduction of the iron ore by carbon (or carbon monoxide) in the blast furnace.

Raw Materials: 1. Iron Ore (Fe_2O_3 or Fe_3O_4) 2. Coke, C
3. Limestone, $CaCO_3$


limestone, coke, iron ore
waste gas to heat exchanger, to heat incoming air
sealing valves
walls of heat-resistant magnesium oxide bricks, cooled by water
hot air
hot air
molten slag
molten iron
From Chemistry for Cambridge IGCSE textbook

- Powdered iron ore, coke and limestone (collectively called the charge) are fed into the top of the blast furnace and heated.
- Hot air is blasted up the furnace from the bottom.
- The coke burns in oxygen from the air to provide heat and produce carbon dioxide:

$$\mathbf{C(s) + O_2(g) \rightarrow CO_2(g) + heat}$$

- Carbon dioxide is reduced by more coke to form carbon monoxide: $\mathbf{CO_2(g) + C(s) \rightarrow 2CO(g)}$
- Carbon monoxide reduce iron (III) oxide to iron:

$$\mathbf{Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)}$$

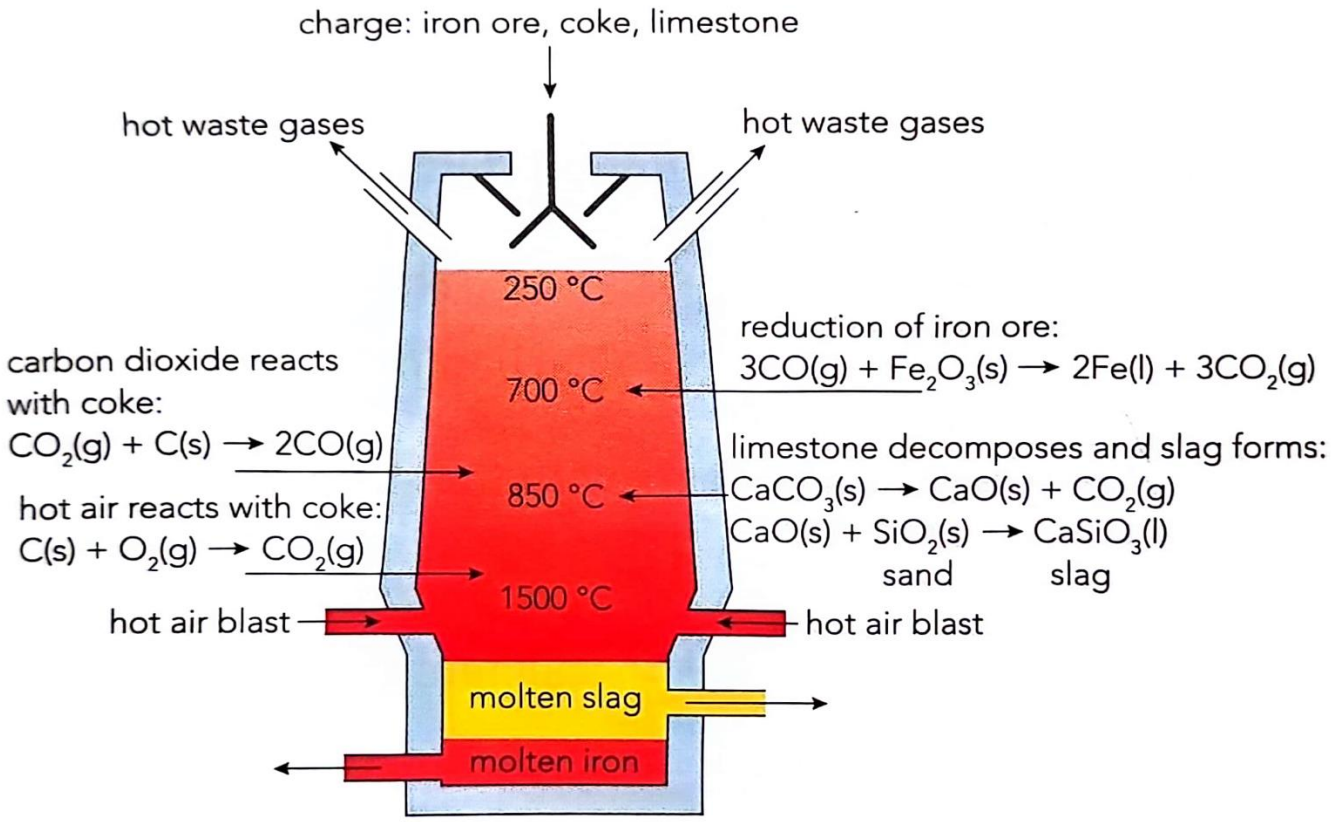
Coke is used instead of coal because it is a better fuel than coal.
It produces more heat and causes less pollution because it contains less amounts of sulfur.

- Dense molten iron runs to the bottom of the furnace and is run off.
  There are many impurities in the iron ore, sand (SiO_2) being the major one. Limestone helps to remove them:
  - Limestone is broken down by heat to calcium oxide: $\mathbf{CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)}$
  - Calcium oxide (a basic oxide) reacts with sand (silicon dioxide, an acidic oxide) to form a liquid called "slag"(calcium silicate $CaSiO_3$): $\mathbf{CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)}$

The liquid slag falls to the bottom of the furnace and floats over the molten iron being less dense, thus it prevents reoxidation of the molten iron.

The slag is tapped off separately. Once solid, it is used in concrete to make buildings and roads.
Hot waste gases escape from the top and are used to transfer heat to the incoming air. This helps reduce the cost of the extraction.

Iron extraction runs continuously and is much cheaper than electrolysis.



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Iron obtained from the blast furnace is around 95% pure and is called **pig iron or cast iron**.

In addition to the element iron, it contains: 4% carbon, sulfur, silicon and phosphorous.
Most of the cast iron is purified further and turned to different types of the alloy steel.

- **Properties of pig or cast iron:**

1. Hard but brittle, it breaks off easily when subjected to pressure
2. Corrodes very easily

- **Uses of cast (pig) iron:**

Pure iron's use is limited due to its ease of rusting in air. Cast iron is usually used in making very large objects like bridges, buildings, ships, …. It is also used in making pots and pans as it conducts heat very well.

CQD30. [G]]Most iron is extracted by reducing iron ore, hematite, by carbon monoxide in a blast furnace.

a. What are the raw materials added to hematite in a blast furnace?
b. A series of reactions occurs in a blast furnace during the extraction of iron from hematite. Describe these reactions.

CQD24.

a. What is cast, or pig iron?
b. What are the major impurities that it may have?

Corrosion and Rusting

Corrosion is the process by which a metal reacts with the **oxygen** of the air in the presence of **water** to produce a layer of oxide. Chemically, corrosion is a redox reaction.

Corrosion of iron or steel to form hydrated iron(III) oxide ($Fe_2O_3.xH_2O(s)$) is called rusting. It is a serious problem and large sums of money is spent annually to replaced iron and steel structures or protecting them. We discussed earlier that aluminum is also attacked by air to produce the oxide, but unlike the iron oxide, the aluminum oxide remains stuck to the metal surface actually protecting it from further oxidation.

CQD23. What two chemicals are essential for rusting to take place?

Rust Prevention

To prevent iron rusting, it is necessary to stop oxygen and water from coming in contact with iron.
This can be achieved by

Barrier methods, such as:

1. **Painting**
   Ships, lorries, cars and many other iron and steel structures are painted to prevent rusting.
   However, if the paint is scratched, the iron beneath will start to rust, that is why the paint must be kept in good condition and checked regularly.

2. **Oiling or Greasing**
   This is the most common way of protecting moving parts of machinery and vehicles, but the protective oil film must be renewed regularly.

3. **Coating with plastic**
   The exteriors of refrigerators, freezers and many other items are coated with plastic such as PVC (polyvinylchloride) to prevent the steel structure from rusting.

Electroplating

a. Cans for food can be made from steel coated with tin. The tin is deposited on to the steel used to make food cans by dipping the steel into molten tin.

b. Galvanizing or Zinc Plating

Galvanizing is the process of covering iron with a thin layer of zinc by either dipping it into molten zinc or electroplating it with zinc.
Galvanization prevents iron from rusting in two ways:

- by keeping water and oxygen away from the iron

- as zinc is more reactive, then it will be oxidized instead of iron.

For example, steel girders used in the construction of building and bridges are galvanized by dipping into molten zinc or electroplating with zinc to protect them from rusting. If the zinc layer is scratched, protection does not stop. As long as some zinc remains in contact with the iron structure, the structure will be protected from corrosion as the more reactive zinc corrodes in place of the iron.

Sacrificial protection (Zn or Mg)

It is the process of attaching a piece of metal (must be more reactive than iron) to steel pipes, hulls of ships or oil rigs to protect them from rusting. The more reactive metal reacts with oxygen and water in the air and sacrifices itself for iron.

Bars of zinc are attached to hulls of ships and to oil rigs to protect them from rusting. When the zinc runs out, it must be renewed. Gas and water pipes made of steel are connected by a wire to blocks of magnesium to obtain the same result

CQD33. [G] List three methods to prevent corrosion of a metal.

CQD39. [G] Aluminum is more reactive than iron, however it is used in many applications such as satellite dishes and ladders without any coating or protection. Explain why.

Uses of Copper

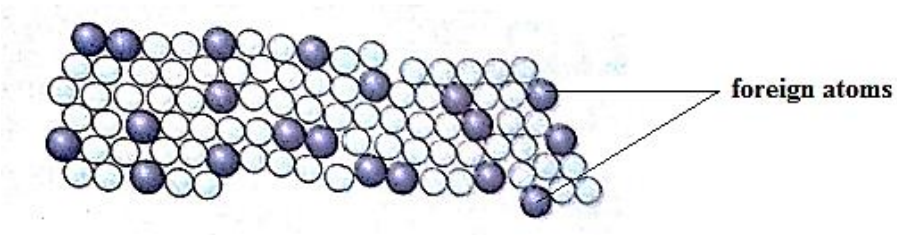
Use	Property
In electrical wires	1. Highly ductile 2. Excellent conductor of electricity
In alloys such as brass (Cu and Zn) or bronze (Cu and Sn) They are used for making statues, machine parts and gears.	1. Hardness 2. Resists corrosion
In steam boilers, cooking utensils and water pipes.	1. Good conductor of heat 2. Resists corrosion

CQD22. List three uses of copper.

Alloys

Alloys are solid solution of two or more metals that may contain non-metals as well. Typical alloys are:

Mixing atoms of only slightly different sizes disrupts the regular packing, and make it much more difficult for layers to slide over each other when a force is applied. This is what makes alloys harder and stronger but more brittle than the individual metals that make them up. Alloys are still metals and so we expect them to exhibit all properties of metals.



It is generally found that alloying produces a metallic substance that has more useful properties than the original pure metals it was made from. For example, brass, an alloy of copper and zinc, is harder and more corrosion resistant than either of the metals it is made from. Of all the alloys we use, steel is perhaps the most important. Many steels have been produced; they contain not only iron but also carbon and other metals. For example, nickel and chromium are the added metals when stainless steel is produced. The chromium prevents the steel from rusting as chromium forms an inert oxide layer at its surface while the nickel makes it harder as it has different atomic size than iron.

Properties of steel:

1. Less brittle than iron
2. More resistant to corrosion than iron.

Types of Steel and their uses:

Type of steel	Percentage composition	Use
Mild steel	0.5 % C, rest is Fe	Car bodies and machinery
Hard steel	1.5 % C, rest is iron	Cutting tools, razors, blades
Stainless steel	Fe, Ni, Cr, C	Chemical plant, cutlery, surgical equipment because of its hardness and resistance to corrosion

High carbon steel are strong but brittle and low carbon steels are softer but more malleable

CQD26. Which of the following statements describing alloys is false?

A. Alloys are harder than the metals they are made from.
B. Alloys are cheaper than the metals they are made from.
C. Alloys are good conductors of electricity.
D. Alloys are mixtures of metals.
E. Alloys can be molded into different shapes.

CQD27.
a. One of the alloys of iron is stainless steel. Stainless steel is more expensive than steel, it resists corrosion and is harder than it. Which of the following is made of steel?

A. Saucepans
B. Cutlery
C. Pipes in a factory
D. Cooking utensils
E. All of the above

b. Cutlery is made of stainless steel while food containers are made of aluminum. What is the property possessed by both metals that make them suitable for the specified use? They are both:

A. good conductors of heat.
B. resistant to corrosion.
C. good conductors of electricity.
D. very strong.
E. molded into different shapes.

CQD31. [G] Describe the structure of alloys.

CQD32. [G] Why are alloys used instead of pure metals? Give an example.

Electroplating

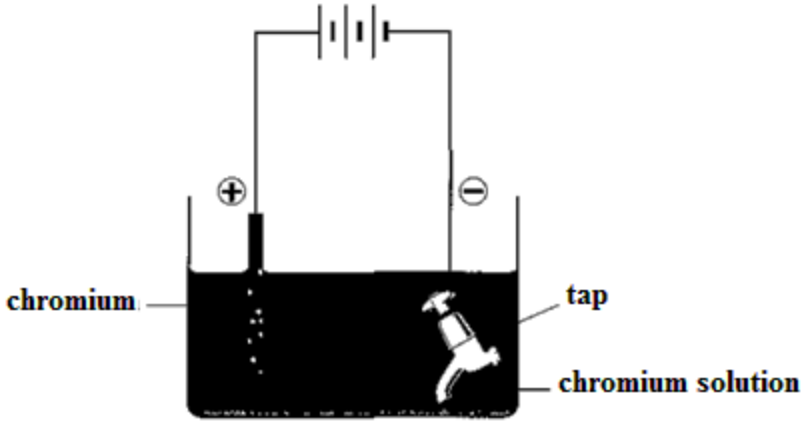
It is the process of covering a metallic object with metal using electricity for more attractive appearance and protection against corrosion.

Steps:

1. Object to be electroplated is made the cathode (connected to the negative terminal of the battery).
2. Metal to electroplate with is made the anode (connected to the positive terminal of the battery).
3. Electrolyte is an aqueous solution of the metal to electroplate with.

- Example: Chromium plating of a tap

Cathode: Tap of water
Anode: Chromium rod
Electrolyte: Chromium nitrate solution



Details to note:

- Clean surface of the object is needed, to ensure that the metal layers stick to the object.
- Rotate the object from time to time, to ensure that the object is uniformly electroplated.
- Dip the object deep in the electrolyte, to ensure that the entire object is electroplated.

CQD28. Draw a labelled sketch of the apparatus that may be used to electroplate an object with silver by electrolysis. Write the equations for the reactions that will take place at each electrode.

Extraction of Aluminium

Aluminium makes up 8.1% by mass of the Earth's crust making it the most abundant metal in the Earth. Aluminum is highly reactive, but its reactivity is masked by a thin oxide layer that coats the metal. Aluminium spontaneously reacts with oxygen of the air to produce aluminium oxide which strongly adheres to the surface of the metal.

Aluminum is the second most widely used metal after iron.

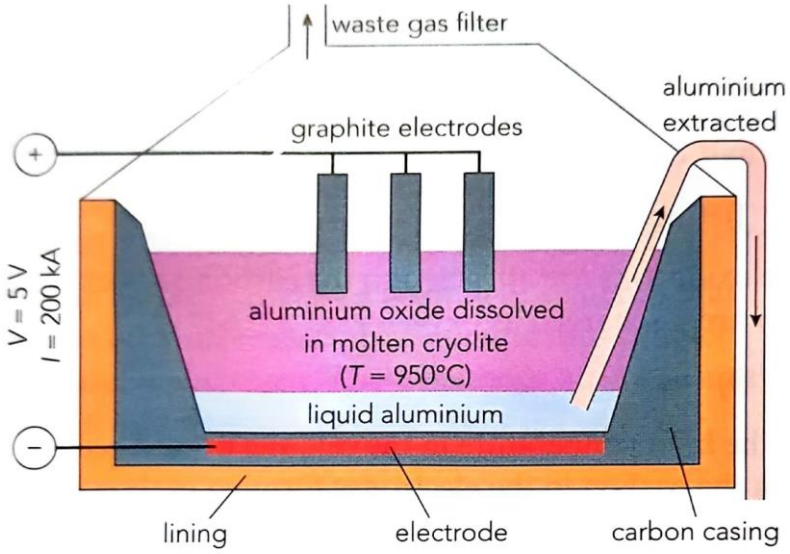
The main ore from which it is extracted is called bauxite. The ore is first mined then purifies and final melted and electrolyzed.

Ore: Bauxite: composed of mainly aluminium mineral, along with iron (III) oxide and other impurities.

Al_2O_3 is amphoteric while Fe_2O_3 is basic. The ore is treated first with sodium hydroxide. The amphoteric Al_2O_3 dissolves in NaOH to form sodium aluminate while the basic Fe_2O_3 does not and is separated by filtration.

Alumina is not soluble and has a melting point of 2017°C. so it is dissolved in molten cryolite Na_3AlF_6 (sodium aluminium fluoride). The cryolite acts as both a solvent to dissolve Al_2O_3 and reduces the melting point of the electrolyte from 2017°C to 800-1000°C. This saves on the energy costs.

The molten liquid is the electrolyzed in a cell using graphite electrodes. The process runs continuously with liquid aluminum removed and more alumina added.



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Reactions:

At anode: $6O^{2-}(l) \rightarrow 3O_2(g) + 12e^-$
At cathode: $4Al^{3+}(l) + 12e^- \rightarrow 4Al(l)$
Overall reaction: $2Al_2O_3(l) \rightarrow 4Al(l) + 3O_2(g)$

Oxygen produced at the anode reacts with it producing CO_2 and CO, so the anode has to be replaced periodically.

Requirements of extraction of aluminium:

- Source of electricity: preferably hydroelectric power.
- *The process uses a great deal of electricity and is not cost-efficient unless the electricity is cheap.*
- Large supply of cooling water.
- Raw materials (Bauxite and cryolite)
- Good transport system.
- Waste gases purification (to get rid of carbon monoxide, CO, and carbon dioxide, CO_2 produced at the anode).

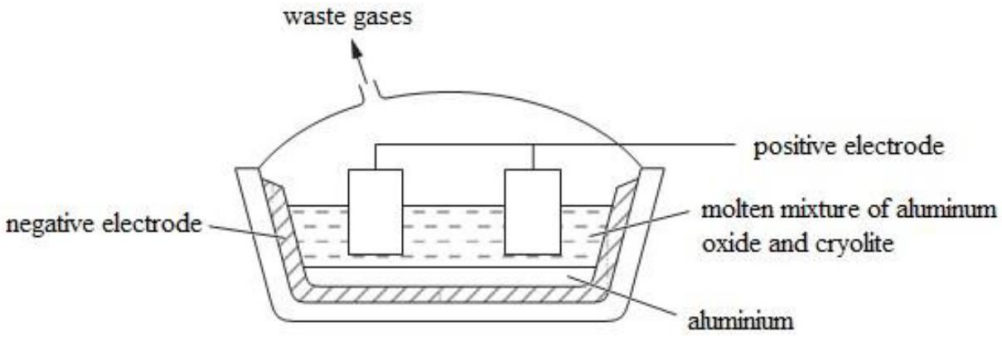
CQD 35.

a. What chemical is used to manufacture aluminum industrially?
b. What are the formula and common name of the aluminum ore?
c. What is added to the aluminum ore to remove the impurity that is present along with it?
d. What is added to the aluminum ore to lower its melting point?
e. What is the function of cryolite?
f. At around what temperature will the process take place?
g. Write an equation to show how aluminum is produced?

CQD36. Aluminum is produced industrially by electrolyzing its oxide.

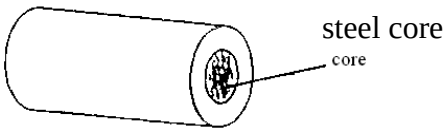
a. What raw material is used at the electrolyte? What other raw material is used to lower the operating temperatures?
b. What type of energy is used in electrolysis?
c. State the materials from which the anode and cathode are usually made.
d. Why does the anode have to be repeatedly replaced?

CQD38. [G] Aluminum is extracted by the electrolysis of molten aluminum oxide (aluminum ore) dissolved in cryolite. A schematic diagram of the apparatus used is shown below.



a. Give two reasons why the electrolysis is done using a molten mixture of aluminum oxide and cryolite instead of solid aluminum oxide.
b. Give the ionic half-equations for the reactions occurring at the electrodes.
c. The anodes made up of graphite have to be replaced regularly. Explain why.

Uses of Aluminium

Use	Property
In cooking utensils In food containers	1. Resists corrosion (due to the protective oxide layer) 2. Low density 3. Good conductor of heat
Making Aircrafts (airplanes)	1. Low density 2. High tensile strength (withstands high pressure)
Steel cored aluminium cable  or Overhead power cables	Aluminium is used because it: **1.** is a good conductor of electricity **2.** has low density (less dense than copper that was previously used) Steel is used because it: 2. expands less than copper 3. has a high melting point. The steel is used to support aluminium which is a very good conductor.

Recycling aluminum requires around 5% of the energy needed to extract it, this is why it is imperative to recycle aluminum.

CQD40. [G] Fill in the table to determine the property related to each given usage of aluminum.

Use	Property
In aircrafts (airplanes)	low density of aluminum
In cooking utensils	excellent thermal conductivity
In electrical wiring systems	excellent electrical conductivity, light weight and malleable.
Food packaging	unbreakable and unreactive oxide
Building industry	strong, light weighted, shiny

Applications

1. The table below gives some information about six different substances

substance	melting point / ˚C	boiling point / ˚C	electrical conductivity as a solid	electrical conductivity as a liquid
A	839	1484	good	good
B	- 210	- 196	poor	poor
C	776	1497	poor	good
D	- 117	78	poor	poor
E	1607	2227	poor	poor
F	- 5	102	poor	good

Choose from A – F to answer the following questions:

a) A substance that has a macromolecular structure like silicon (IV) oxide
b) 3 substances that are solids at room temperature
c) A substance that is a metal
d) A substance that could be **aqueous** potassium chloride
e) A substance that is an ionic compound
f) 2 substances that are liquids at room temperature

2. Choose an element from the list below to answer the following questions: **Rb, Fe, Si, P, I, Sr**

a) An element that reacts with cold water
b) An element that is a solid at room temperature and exists as diatomic molecules.
c) An element that can form the oxides XO and X_2O_3.
d) An element that has the hydride of the type XH_3.
e) An element that has a macromolecular structure similar to that of carbon.

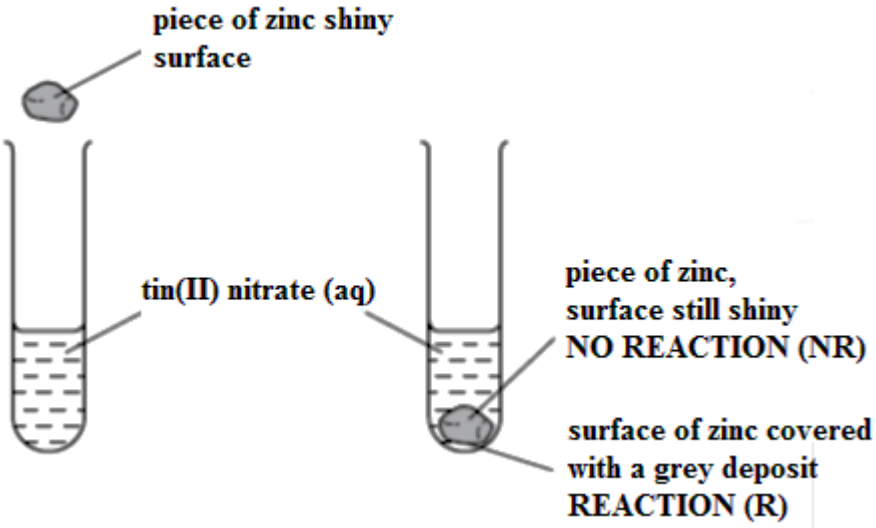
3. Tin is a Group IV element, its position of element tin in the reactivity series is as follows

Zinc
Iron
Tin
Copper

Complete the following reactions (if any), or else write "no reaction"

$Fe + Sn^{2+} \rightarrow$
$Sn + Zn^{2+} \rightarrow$
$Cu + Sn^{2+} \rightarrow$

4. The reactivity series gives the order of reactivity of metals
An experiment below was done to find out which is a more reactive metal, tin or zinc.


piece of zinc shiny surface
tin(II) nitrate (aq)
piece of zinc, surface still shiny NO REACTION (NR)
surface of zinc covered with a grey deposit REACTION (R)

The above experiment can be done with other metals and the results could be recorded in a table, then the order of reactivity can be predicted.

i) The order was found to be as follows:

manganese	**most reactive**
zinc	
tin	
silver	**least reactive**

Complete the table below

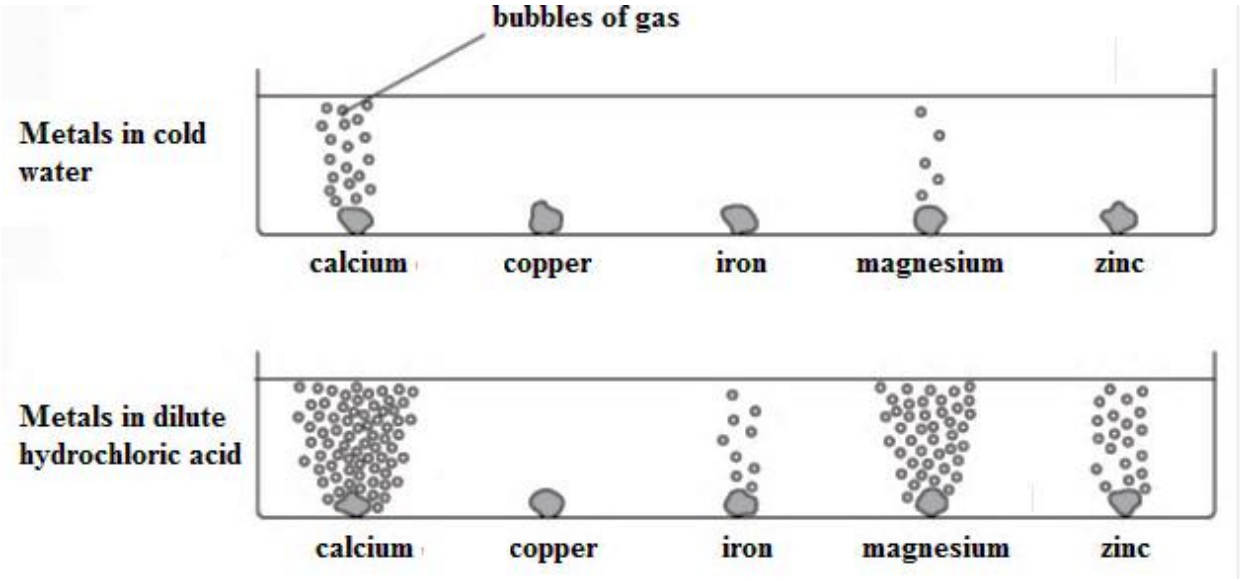
aqueous solution	tin Sn	manganese Mn	silver Ag	zinc Zn
tin(II) nitrate		R	NR	R
manganese(II) nitrate				
silver(I) nitrate				
zinc nitrate				

ii) Write an ionic equation for the reaction between silver ions and tin atoms.

iii) The reaction below is a redox reaction: $Mn + Sn^{2+} \rightarrow Mn^{2+} + Sn$
Which change is oxidation? Explain

iv) Explain why the above experiment cannot be used to predict the position of aluminium in the reactivity series.

5. The diagrams below show the reactions of some metals with cold water and with dilute hydrochloric acid.



Answer the following questions, using only the metals that appear in the diagrams.

a) Name two metals that react with cold water.
b) Name one metal that reacts with dilute hydrochloric acid but not with cold water.
c) Arrange the five metals in order of reactivity.

6. Some reactions of metals **W**, **X**, **Y** and **Z** are given below.

Metals	Reaction with water	Reaction with dilute hydrochloric acid
W	A few bubbles form slowly in cold water	Vigorous reaction. Gas given off.
X	Vigorous reaction. Metal melts. Gas given off.	Explosive reaction. Should not be attempted
Y	No reaction.	No reaction.
Z	Does not react with cold water. Hot metal reacts with steam.	Steady fizzing.

a) Arrange these metals in order of reactivity.
b) Which of these metals could be: magnesium, copper
c) The equation for the reaction of X with cold water is given: $2X(s) + 2H_2O(l) \rightarrow 2XOH(aq) + H_2(g)$
   (i) Describe the test you would use to show that the gas evolved is hydrogen.
   (ii) How could you show that the water contained a compound of the type XOH?
   (iii) In which group of the Periodic Table does metal X belong?
   (iv) The ore of X is its chloride. Suggest how metal X could be extracted from its chloride.

7. The **reactivity series** of some metals is given below. Unfamiliar elements are marked by (*) and their common oxidation numbers are given.

*Barium	Ba
*lanthanum	La (+3)
Magnesium	
Zinc	
*chromium	Cr (+2), (+3), (+6)
iron	
copper	
*palladium	(+2)

Choose from the above metals to answer the following questions

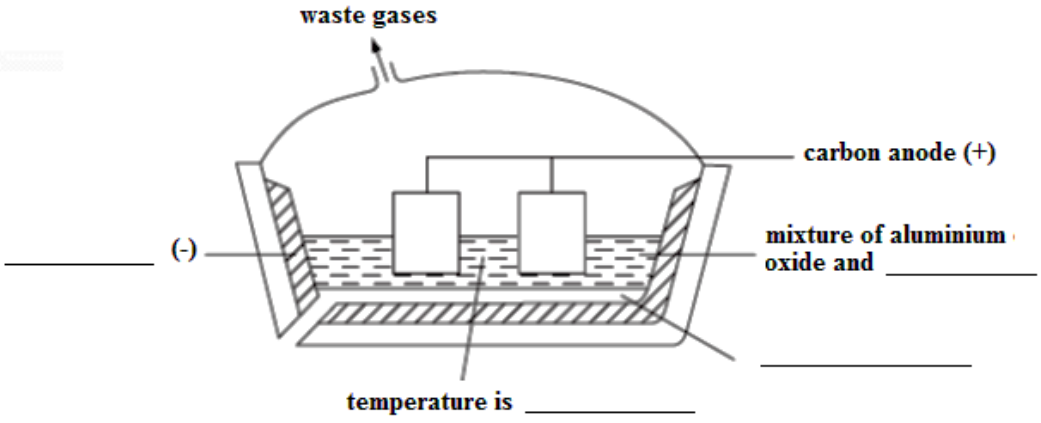
a) Metals that cannot react with dilute hydrochloric are …… and ……..
b) The unfamiliar metals that can react with cold water are ……..and ……..
c) The oxidation state of barium is …….
d) The unfamiliar metal whose oxide cannot be reduced by carbon is …….
e) Explain why is it possible to predict why metals like chromium and iron have more than one oxidation state.

8. Aluminium is extracted by the electrolysis of a molten mixture that contains alumina (aluminium oxide, Al_2O_3)

a) Bauxite is an aluminium ore. This ore contains alumina, that is amphoteric, and iron (III) oxide, that is basic. The ore is heated with aqueous sodium hydroxide. Complete the following:

The …….. dissolves to give a solution of ……..….
The …….. does not dissolve and can be removed by …….

b) Complete the labeling of the following diagram.



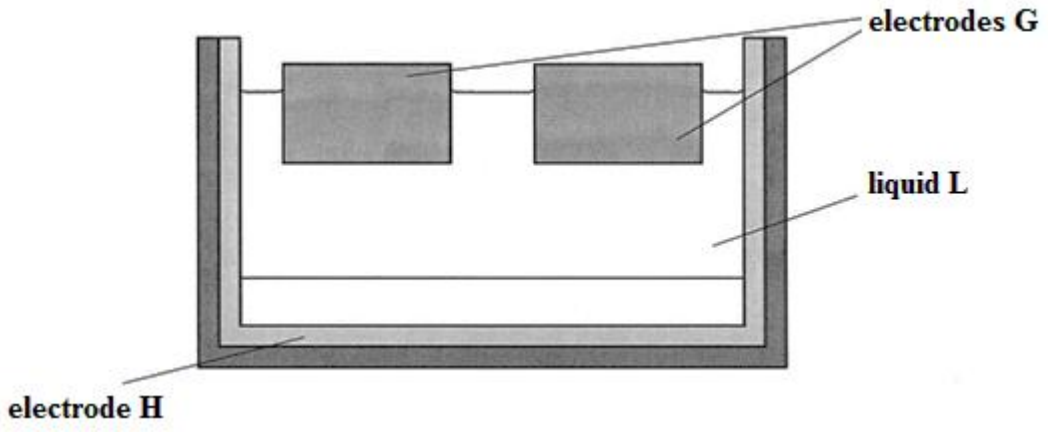
c) The ions involved in electrolysis are Al^{3+} and O^{2-}

i. Write the equation for the reactions that occur at anode and cathode.
ii. How is carbon dioxide formed at the anode?

9. Aluminium was first extracted using sodium in 1827. $AlCl_3 + 3Na \rightarrow Al + 3NaCl$
Aluminium, obtained by this method, was found to be more expensive than gold.

a) Explain why aluminium was so expensive.
b) The modern method for the extraction of aluminium is by the electrolysis of the molten electrolyte, aluminium oxide, dissolved in cryolite where the aluminium oxide decomposes according to the reaction: $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2$
Both electrodes used are made of carbon. The oxide is dissolved in cryolite. Give two reasons for this.
c) Why do the carbon anodes need to be replaced frequently?
d) The modern method of extraction by electrolysis of the molten electrolyte is one method of extraction. Other methods include the electrolysis of the aqueous solution and the reduction of the oxide by carbon. Explain why the last two methods cannot be used to extract aluminium.
   1) Electrolysis of an aqueous solution
   2) Using carbon
e) Explain the following
   i. Aluminium is used extensively in the manufacture of aircraft.
   ii. Aluminium is used to make food containers.
   iii. Aluminium electricity cables have a steel core.

10. The diagram below represents how aluminium is extracted industrially.



a)
(i) The process used to extract aluminium is ……
(ii) The element used to make electrodes G is …….
(iii) Electrode H is the ……. (positive/negative) electrode.
(iv) Liquid L contains aluminium oxide with another substance. What is this substance and why is it used?

b) The product produced at electrode G reacts with the electrode to form carbon monoxide and carbon dioxide.
(i) This product is …….
(ii) Explain why carbon monoxide is poisonous.

c) Aluminium has many uses that depend on its structure and the physical properties.
(i) The strength of aluminium depends on the electrostatic force of attraction between two types particle present in its structure.
Give the name of these two types of particles

(ii) Aluminium is ductile because it can be easily pulled in to a wire. Explain, using structure, why aluminium is ductile.

(iii) Aluminium is used in manufacturing aircraft bodies. State the property that makes it suitable for this use.

11. Aluminium ore is called bauxite that is impure aluminium oxide. Pure aluminium oxide, alumina, is obtained from bauxite.

Aluminium is produced at the cathode when a molten mixture of alumina and cryolite, Na_3AlF_6, is electrolyzed.

a) The major impurity in bauxite is iron (III) oxide. Aluminium oxide is amphoteric while iron (III) oxide is basic. Explain how a solution of sodium hydroxide can be used to separate them.

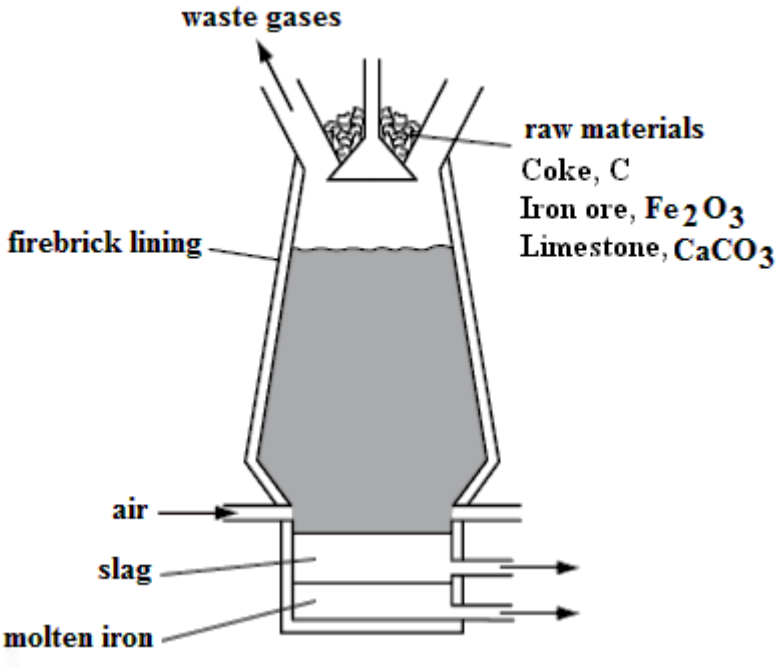
b) The process of purification of bauxite uses large amounts of sodium hydroxide.

i. Describe the chemistry of how sodium hydroxide is made from concentrated aqueous sodium chloride. Your description must include at least one ionic equation.

ii. The production of sodium hydroxide from sodium chloride produces two other chemicals. Complete the table below:

Chemical	Use

12. Iron is extracted using the blast furnace from its ore heamatite using reduction.



a) What is meant by the term reduction?
b) What is the chemical name of heamatite?
c) The main heat source in the furnace is via the process of burning coke in air. $C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$
The reaction that produces heat is called ……
d) Write the equation to show how carbon monoxide, reducing agent in the furnace, is produced.
e) Balance the following equation: $Fe_2O_3(s) + CO(g) \rightarrow Fe(l) + CO_2(g)$
f) Limestone is added to the blast furnace to help in the removal of impurities in the ore such as silicon dioxide. Explain the chemistry of this step.
g) Write the equation for the formation of slag.
h) Iron from the blast furnace is impure. Two of the impurities are silicon and carbon that are removed by blowing oxygen through molten iron and adding calcium oxide.
   i. How does the addition of oxygen remove carbon?
   ii. How does the addition of oxygen and calcium oxide remove silicon?

13. Iron is extracted in a blast furnace. The list below gives some of the substances used or formed in the extraction: **carbon monoxide coke iron ore limestone slag**

a) Which substance is a mineral containing largely calcium carbonate?
b) Which substance is formed when impurities in the ore react with calcium oxide?
c) Which substance is also called hematite?
d) State two functions of the coke used in the blast furnace.
e) Most of the iron is converted into mild steel or stainless steel. Give one use mild steel and one for stainless steel.

14. Iron and aluminium are two important metals extracted from their ores on a large scale.

a) In the extraction of iron, three different raw materials are put into the top of a blast furnace. Name the main compound present in the following raw materials: haematite and limestone

b) The following equations represent reactions in the blast furnace.

A $C + O_2 \rightarrow CO_2$
B $CaCO_3 \rightarrow CaO + CO_2$
C $C + CO_2 \rightarrow 2CO$
D $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
E $CaO + SiO_2 \rightarrow CaSiO_3$

Choose from the letters **A**, **B**, **C**, **D** or **E** to answer parts (i) – (iv). Each letter may be used once, more than once or not at all.

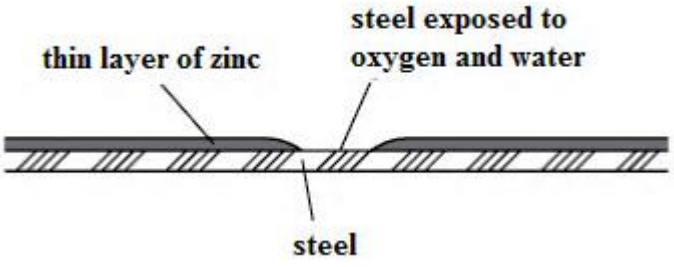
(i) A reaction that is used to produce heat
(ii) A neutralisation reaction
(iii) A decomposition reaction
(iv) A reaction that forms a reducing agent

c) Molten iron and another molten substance collect at the bottom of the blast furnace. What is the common name of this other molten substance?
d) Aluminium is extracted from its ore by electrolysis. This is a more expensive process than using a blast furnace.
   (i) Why is a different method used for aluminium?
   (ii) State the major reason for the high cost of extracting aluminium.
e) Coke used in the blast furnace contains carbon. Carbon is also used in the extraction of aluminium, but for a different purpose. What is this purpose?

15. Steel is an alloy manufactured from impure iron.
Iron and steel both rust. The formula of rust is $Fe_2O_3.2H_2O$ (hydrated iron (III) oxide).

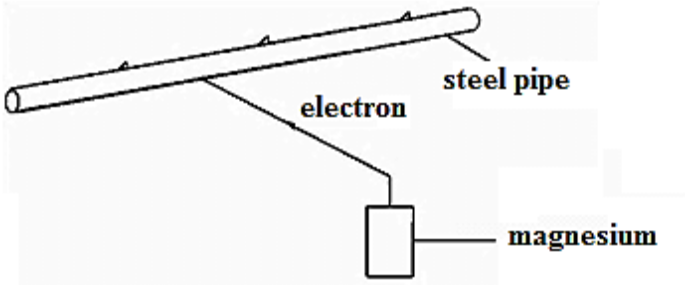
a) The substances that must be present for rusting to occur are …… and …….
b) Name two methods, other than painting and coating with grease, that are used for preventing rusting of iron and steel.

16. The major use of zinc is in galvanizing where steel objects are coated with a thin layer of zinc. This process protects steel from rusting even if the layer of zinc is broken.

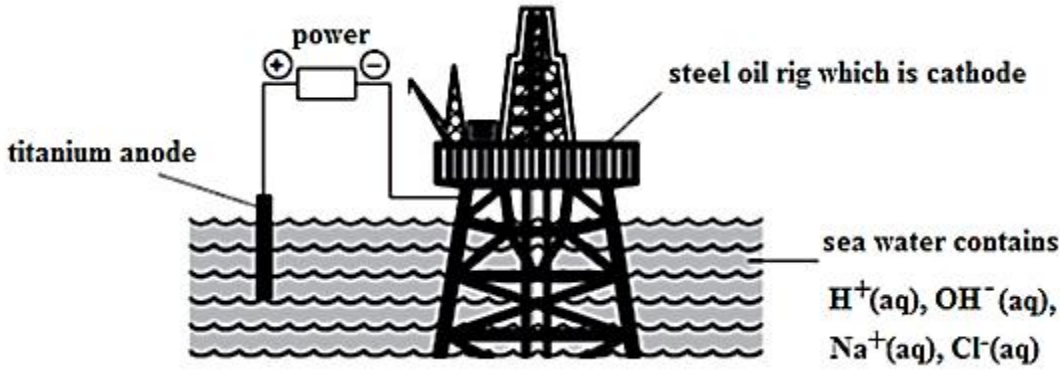


Explain, using ions and electrons, why the steel that is exposed does not rust.

17. Explain why attaching blocks of magnesium to steel pipelines prevents them from rusting.



18. Titanium metal is highly resistant to corrosion. One of the uses of titanium is in the cathodic protection of large structures of steel from rusting.



a) Define oxidation in terms of electron transfer.
b) Name the gas produced at the steel oil rig that is the cathode
c) The gases formed at the titanium anode are …… and ……
d) The oil rig does not rust. Explain
e) Another method of protecting steel from corrosion is by sacrificial protection. State two differences between cathodic protection and sacrificial protection.

19. Iron is a useful metal. One problem with using iron is that it can rust.

a)
i) Name the iron compound present in rust.
ii) Name the **two** substances that iron reacts with when it rusts.

b) What type of reaction occurs in the rusting of iron?

c) Galvanising can prevent iron from rusting. In this process, the iron is coated with another metal.
i) Identify the other metal.
ii) Identify one object suitable for galvanizing: bicycle chain, **bucket**, car engine, drink can

20. Steel may be coated with another metal as zinc or chromium or with a polymer as poly(chloroethene) to prevent rusting.

a) Suggest a property of poly(chloroethene) that makes it suitable for this purpose.
b) Explain why the steel will rust when the protective coating of polymer is broken?

Steel may be coated with another metal as zinc or chromium or with a polymer as poly(chloroethene) to prevent rusting.

c) Suggest a property of poly(chloroethene) that makes it suitable for this purpose.
d) Explain why the steel will rust when the protective coating of polymer is broken?

21. Iron is a transition element.

a) Which of the following statements about transition elements are correct?

- The metals are highly coloured e.g. yellow, green, blue.
- The metals have low melting points.
- Their compounds are highly coloured.
- Their compounds are colourless.
- The elements and their compounds are often used as catalysts.
- They have more than one oxidation state.

b)

i) In which Period in the Periodic Table is iron to be found?

ii) Use the Periodic Table to work out the number of protons and the number of neutrons in one atom of iron.

22. Vanadium is a transition element.

a) Predict two differences between the physical properties of vanadium and potassium.

b) Predict two differences between the chemical properties of vanadium and potassium.

Essay Questions:

23. Metal pendants can be coated with a layer of copper using electricity. Plan an experiment to electroplate a small metal pendant with copper. You are provided with common laboratory apparatus, a copper rod, copper(II) sulfate crystals, water and a small metal ring. Include a labeled diagram in your answer.

Topic 11-Qualitative Chemical Analysis

Species tested	Test	Result & Equation
<u>Test for Cations</u>		
Li^+ Lithium ion	Flame Test	Red flame
K^+ Potassium ion	Flame Test	Lilac flame
Na^+ Sodium ion	Flame test	Yellow flame
Ca^{2+} Calcium ion	Flame test	Brick red flame
Cu^{2+} Copper (II) ion	Flame test	Blue-green flame
Ba^{2+} Barium ions	Flame test	Pale green
Flame test is used to show the presence of a certain metal ion in a compound: • **A platinum or nichrome wire is cleaned by dipping in concentrated hydrochloric acid and then holding it in a hot Bunsen flame.** • **This is repeated until the wire does not give any colour in the flame.** • **The wire is dipped back into the concentrated acid, then into a tiny sample of the solid you are testing and back into the flame.**		
Species tested	**Reagent used**	**Result & Equation**
NH_4^+ Ammonium ion	An alkali, for eg. NaOH(aq) add to an aqueous solution of the specimen to be tested then warm gently	No precipitate but a pungent smelling gas is produced that turns damp red litmus paper blue. $NH_4^+(aq) + OH^-(aq) \xrightarrow{\Delta} NH_3(g) + H_2O(l)$
Ca^{2+} Calcium ion	NaOH(aq) add to an aqueous solution of the specimen tested	A white precipitate is formed that does **<u>not</u>** dissolve in excess NaOH(aq) $Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$
	$NH_3(aq)$ add to an aqueous solution of the specimen tested	No reaction / no precipitate

Species tested	Reagent used	Result & Equation
Al^{3+} Aluminium ion	NaOH(aq) add to an aqueous solution of the specimen tested	A white precipitate is formed that dissolves in excess NaOH(aq) $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$
	$NH_3(aq)$ add to an aqueous solution of the specimen tested	A white precipitate of $Al(OH)_3$ is formed that does **not** dissolve in excess $NH_3(aq)$. $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$
Zn^{2+} Zinc ion	NaOH(aq) add to an aqueous solution of the specimen tested $NH_3(aq)$ add to an aqueous solution of the specimen tested	**In both cases:** A white precipitate of $Zn(OH)_2$ is formed that dissolves in excess reagent $Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$
Fe^{2+} Iron (II) ion	NaOH(aq) add to an aqueous solution of the specimen tested $NH_3(aq)$ add to an aqueous solution of the specimen tested	**In both cases:** Green precipitate is formed that does **not** dissolve in excess reagent $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
Fe^{3+} Iron (III) ion	NaOH(aq) add to an aqueous solution of the specimen tested $NH_3(aq)$ add to an aqueous solution of the specimen tested	**In both cases:** Red-brown precipitate is formed that does **not** dissolve in excess reagent $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

Species tested	Reagent used	Result & Equation
Cu^{2+} Copper (II) ion	NaOH(aq) add to an aqueous solution of the specimen tested	Blue precipitate of is formed that does **<u>not</u>** dissolve in excess NaOH(aq) $Cu^{2+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$
	$NH_3(aq)$ add to an aqueous solution of the specimen tested	Blue precipitate of $Cu(OH)_2$ is formed that dissolves in excess $NH_3(aq)$ to form a deep blue solution. $Cu^{2+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$
Cr^{3+} Chromium(III) ion	NaOH(aq) add to an aqueous solution of the specimen tested	Grey green precipitate is formed that dissolves in excess NaOH(aq) $Cr^{3+}(aq) + 3OH^{-}(aq) \rightarrow Cr(OH)_3(s)$
	$NH_3(aq)$ add to an aqueous solution of the specimen tested	Grey-green precipitate that does **<u>not</u>** dissolve in excess $NH_3(aq)$ $Cr^{3+}(aq) + 3OH^{-}(aq) \rightarrow Cr(OH)_3(s)$

Test for Anions		
Species tested	**Test**	**Result & Equation**
CO_3^{2-} Carbonate ion	Dilute acid (HCl(aq) or HNO_3(aq)) Add dilute acid to a sample of the specimen to be tested (solid or aqueous)	Effervescence/bubbles or fizzing is observed. A colorless, odorless gas is produced that turns lime water milky when passed through it for a short time. $MCO_3(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + H_2O(l) + CO_2(g)$ OR $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$ If the gas is passed through lime water for a **long time** the milkiness dissolves and the solution becomes clear again.
SO_4^{2-} Sulphate ion	$Ba(NO_3)_2$(aq) add to an aqueous solution of the specimen to be tested followed by a dilute acid (HCl(aq) or HNO_3(aq)) OR Acidified $Ba(NO_3)_2$(aq) add to an aqueous solution of the specimen to be tested	White precipitate is formed that does **not** dissolve in acid. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ **The solution is acidified to make sure there is no carbonate ion CO_3^{2-} or sulphite ion SO_3^{2-}.** **If a carbonate ion CO_3^{2-} or sulphite ion SO_3^{2-} is present, they will not precipitate in acidified reagent.**

Species tested	Reagent used	Result & Equation
SO_3^{2-} Sulphite ion	Add dilute hydrochloric or nitric acid to a solid or aqueous sample then warm gently OR Add acidified potassium manganate(VII)	A colorless, chocking gas is formed that will decolorize purple acidified aqueous potassium manganate(VII) $MSO_3(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + H_2O(l) + SO_2(g)$ OR $SO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + SO_2(g)$ Purple solution decolorized
Species tested	**Test**	**Result & Equation**
NO_3^- Nitrate ion	Aluminium + NaOH(aq) add to an aqueous solution of the specimen to be tested then warm gently	A pungent smelling gas is produced that turns damp red litmus paper blue. (Ammonia gas)
Cl^- Chloride ion	$AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested followed by dilute acid ($HCl(aq)$ or $HNO_3(aq)$) OR Acidified $AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested	A white precipitate is formed, which upon exposure to light turns violet then black. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Species tested	Reagent used	Result & Equation
Br^- Bromide ion	$AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested followed by dilute acid (HCl(aq) or $HNO_3(aq)$) OR Acidified $AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested	A pale cream precipitate is formed, which upon exposure to light turns violet then black. . $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$
I^- Iodide ion	$AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested followed by dilute acid (HCl(aq) or $HNO_3(aq)$) OR Acidified $AgNO_3(aq)$ add to an aqueous solution of the specimen to be tested	A yellow precipitate is formed, which upon exposure to light turns violet then black. $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$
The acid is added to exclude other substances that might also produce precipitates with silver nitrate solution.		

Test for Gases		
H_2 Hydrogen gas	Lit splint	Explodes (burns with a squeaky popping sound)
O_2 Oxygen gas	Glowing splint	Glowing splint relights
NH_3 Ammonia gas	Damp red litmus paper	Turns blue.
CO_2 Carbon dioxide gas	Lime water	Pass the gas through lime water for a **short time**, lime water turns cloudy white (milky). $Ca^{2+}(aq) + 2OH^-(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$ If the gas is passed through lime water for a **long time** the precipitate dissolves and the solution becomes clear again.
Cl_2 Chlorine gas	Damp litmus paper	blue litmus paper turns red then gets bleached red litmus paper gets bleached
SO_2 Sulfur dioxide	acidified aqueous potassium manganate (VII)	Decolorizes the purple solution
H_2O Water vapour **OR liquid**	anhydrous copper(II) sulfate	Turns blue cobalt chloride paper pink $CuSO_4(s) + 5H_2O(l/g) \rightarrow CuSO_4.5H_2O(s)$ white blue
	anhydrous cobalt(II) chloride paper	Turns white anhydrous copper(II) sulfate blue $CoCl_2(s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O(s)$ blue pink

Cobalt chloride paper is a filter paper that has been dipped into cobalt (II) chloride solution then dried thoroughly in a dessicator.
A dessicator is a piece of glassware or a small cabinet which contains a tray of some substance that absorbs water.

Properties of Gases:

Gas	Colour	Odour	Method of collection
$H_2(g)$ Hydrogen	Colourless	Odourless	Less dense than air, collected by: • Upward delivery • Over water (insoluble) • Using a gas syringe
$O_2(g)$ Oxygen	Colourless	Odourless	Density close to that of air, collected: • Over water (insoluble) • Using a gas syringe
$CO_2(g)$ Carbon dioxide	Colourless	Odourless	Denser than air, collected by: • Downward delivery • Using a gas syringe • Over water (insoluble)
$Cl_2(g)$ Chlorine	Pale green	Has a pungent smell	Denser than air, soluble, collected by: • Downward delivery • Using a gas syringe
$SO_2(g)$ Sulphur dioxide	Colourless	Smells like burnt matches	Denser than air, collected by: • Downward delivery • Using a gas syringe
$NH_3(g)$ Ammonia	Colourless	Has a pungent smell	Less dense than air, by: • Upward delivery • Using a gas syringe

CQD 1.
a. How can you verify the presence of H^+ ions in a solution using a pH meter?

b. How can you verify the presence of H^+ ions in a solution using the materials listed below? Write the net ionic equation for each reaction and report any expected observations.

i. Sodium carbonate
ii. A piece of magnesium ribbon

c. How are the ions of lithium, sodium, and potassium identified?

d. A flame test is conducted on a sample of calcium chloride. What color will the flame have?

e. How can you test for calcium ions using a concentrated solution of sodium hydroxide? Write the net ionic equation for the reaction and report the expected observation.

CQD 2. How can you test for the following ions using sodium hydroxide and ammonium solutions? Write the net ionic equation for the reaction between the ions tested and the alkali used, and report any expected observations.

a. Al^{3+}
b. Zn^{2+}

CQD 3.
a. How can you test for the following ions using sodium hydroxide solution? Write the net ionic equation for the reaction, and report the expected observations.

i. Fe^{2+}
ii. Fe^{3+}
iii. Cu^{2+}
iv. Cr^{3+}

b. How can you test for copper(II) ions using ammonia solution? Report the expected observations. (No equations are necessary.)

CQD 4. How can you identify ammonium ions using a solution of sodium hydroxide?
Write the net ionic equation for the reaction, and report any expected observations.

CQD 5.
a. How can you verify the presence of OH^- ions in solution using a pH meter?

b. How can you test for the following ions using silver nitrate solution? Write the net ionic equations for the reactions and report the expected observations.

i. Cl^-
ii. Br^-
iii. I^-

CQD 6. How can you test for the following ions using hydrochloric acid solution? Write the net ionic equation for the reaction, and report the expected observations.

i. CO_3^{2-}
ii. HCO_3^-
iii. SO_3^{2-}

CQD 7.
a. How can you test for SO_4^{2-} ions in a solution? Write the net ionic equation for the reaction, and report the expected observations.

b. Write the net ionic equation for the reaction that will take place when a few drops of sodium sulfide, Na_2S, are added to a solution of lead(II) nitrate, $Pb(NO_3)_2(aq)$. What observations are expected in this case?

CQD 8. Describe how each of the following gases are tested. Write an equation for expected reactions (where applicable).

	Test	**Observation (Reaction where applicable)**
H_2 Hydrogen gas		
O_2 Oxygen gas		
NH_3 Ammonia gas		
CO_2 Carbon dioxide gas		
Cl_2 Chlorine gas		
SO_2 Sulfur dioxide		

CQD 9. [G] Identify the unknown gases from the given observations.

a. A yellow-green gas with a chocking smell that bleached a moist litmus paper.
b. A colorless odorless gas that lit a glowing splint.
c. A colorless odorless gas that made a 'pop' sound when approached by a lit splint.
d. A colorless gas with a pungent smell that turned moist litmus paper blue.
e. A colorless gas that turns lime water milky. f . A colorless gas that decolorizes a purple acidified aqueous potassium manganate(VII).

CQD 10. How can you test for the presence of water in a sample of colorless liquid? Write an equation for the expected reaction.

CQD 11. How can you distinguish between hexane and hexene both of which are colorless liquids at room temperature? Do not write equations for any expected reactions.

CQD 12. [T] A student is asked to identify two salts **T** and **W**.

a. The student noted the following observations while testing salt **T**.

Observation 1: solution of **T** forms a white precipitate with aqueous sodium hydroxide and the precipitate does not dissolve upon adding excess of sodium hydroxide.

Observation 2: solution of **T** forms a yellow precipitate with acidified aqueous silver nitrate.

i. What is the cation of salt **T**?
ii. What is the anion of salt **T**?
iii. Deduce the formula of salt **T**.

b. The student noted the following observations while testing salt **W**.

Observation 1: salt **W** gives a red color in a flame test

Observation 2: solution of **W** releases a gas that turns blue wet litmus red upon addition of dilute hydrochloric acid and gentle heating.

i. What is the cation of salt **W**?
ii. What is the anion of salt **W**?
iii. Deduce the formula of salt **W**.
iv. Describe how a flame test is done on a sample of salt **W**.

CQD 13. [T] A student was analyzing two solids **A** and **B**.

Solid **A** is zinc iodide but solid **B** is unknown. a. Tests on solid **A**.

Solid **A** was dissolved in water and the resulting solution was divided into three test tubes in equal portions.

Complete the expected observations:

Portion	Test performed	Observations
First portion	A few drops of aqueous sodium hydroxide were added	
	An excess of aqueous sodium hydroxide was then added to this mixture	
Second portion	A few drops of aqueous ammonia were added	
	An excess of aqueous ammonia was then added to this mixture	
Third portion	dilute nitric acid and aqueous silver nitrate were added to the solution.	

b. Tests on solid **B**
After dissolving solid **B** in water, the following observations were recorded while testing the solution **B**.

Test on solution of **B**	Observations
dilute solution of sodium hydroxide was added	Pale blue precipitate formed
Aluminum foil and aqueous sodium hydroxide were added, then the mixture was heated and the gas produced was tested with litmus paper	Bubbles released that turns wet litmus blue

Write the chemical formula for solid **B**

CQD 14. [T] A student analyzed soluble carbonate salt **R**.

a. No observations were recorded when the carbonate salt **R** was slightly heated in a test tube. What does this test indicate about salt **R**?

b. Salt **R** was dissolved in water and divided equally into four test tubes.
The following tests were conducted on the test tubes. Complete the observations for each of the following tests.

test 1: a small amount of dilute of hydrochloric acid was added to the first test tube. The mixture effervesces producing a gas.

test 2: The gas released was bubbled in lime water solution.

i. Observation:
ii. Identify the gas released.

c. Two more tests were conducted and the following observations were recorded.

tests on solution of salt **R**	observations
test 3: To a new test-tube of solution, dilute sodium hydroxide was added. The mixture was then warmed carefully.	effervescence pungent gas given off gas turned damp litmus paper blue

i. Identify the ion present in the solution from the observation of **test 3**?
ii. Why did the damp litmus turn blue?
iii. Identify the gas released.

d. Write the chemical formula for salt **R**.

e. Why it's necessary to warm carefully the mixture in **test 3**? Explain your answer.